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ABSTRACT

ChemSource is designed as a strategy to help preservice and inservice high school chemistry teachers promote student learning more effectively. Its major premise is that well-designed laboratory investigations are an important avenue for cultivating student interest, engagement, and meaningful learning in chemistry. The SourceBook component of ChemSource is a resource containing 36 content modules and additional supporting materials which aim at providing specific teaching tips and the best available instructional ideas and information gleaned from experienced, successful chemistry teachers. Each content module features: Content-in-a-Nutshell; Laboratory Activities; Demonstrations; Common Misconceptions; Humor; Applications and Implications; Analogies; Reference Resources; History; and Media. General materials on pedagogy, classroom management, safety, teaching chemistry to the disabled, and other useful resources and tabulations precede and follow the content modules. Volume 1 includes the following: Chemical Pedagogy; User's Guide; General Resources; Acids and Bases; Alkali Metals; Atomic Structure; Biogeochemical Cycles; Chemical Bonding; Condensed States; Electrochemistry; and Enzymes. (JRH)

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SourceBook

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TO THE EDUCATIONAL RESOURCES INFORMATION CENTER (ERIC)

VOLUME 1

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

SE 055 640

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Introductory: General Resource (INTR)
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Alkali Metals (ALKA)
Atomic Structure (ATOM)
Biogeochemical Cycles (BICY)
Chemical Bonding (BOND)
Condensed States (COND)
Electrochemistry (ELEC)
Enzymes: Biochemical Catalysts (ENZY)

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Subject Index (INDS)

A SourceBook Module

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

PREFACE AND ACKNOWLEDGEMENTS

Preface

PREFACE

ChemSource is designed as a strategy to help pre-service and inservice high school chemistry teachers promote student learning more effectively. Its major targets are pre-service chemistry teachers, crossover chemistry teachers (teachers educated in an area of concentration other than chemistry) and teachers in their initial years of chemistry teaching, although many experienced chemistry teachers will find the information gathered here helpful. Its major thrust and premise is that well-designed laboratory investigations are an important avenue for cultivating student interest, engagement, and meaningful learning in chemistry. When students can see, hear, touch, smell, and observe the concrete world about them, their appreciation of chemistry as the science of matter and its changes is greatly enhanced.

The Sou*rceBook* component of *ChemSource* is a resource containing thirty-six content modules and additional supporting materials. It is, as the name implies, a resource. It is not meant as a replacement or substitute for textbooks and curricula, but as a means to supplement, enhance, and support existing curricula. Many *SourceBook* activities have been published elsewhere, and have been tested and successfully used by many chemistry teachers. Still others are new ideas contributed by members of the *SourceBook* writing teams and by numerous teachers throughout the United States. "Something old, something new, something borrowed, and all true " is, we think, an accurate description of the contents of *SourceBook*. We are confident that there is something new for everybody within these covers.

SourceBook materials were developed by twelve writing teams, each consisting of three members - two classroom teachers and one college or university teacher. These teams met and discussed format and content for the modules several times during the course of development. They sifted existing literature and combined and gathered what we believe are the best educational strategies, ideas and practice. The teams included good ideas of the past, the best of present practice, and gave each module a future thrust by incorporating new developments in chemical research in a number of places in the modules. The way that *SourceBook* was compiled is an excellent illustration of the value of the strategy of "teachers teaching teachers."

On the following pages, we acknowledge the many writers, support staff, laboratory activity developers, contributors of ideas, links, connections, strategies, and numerous others who have, in various ways, helped to bring these volumes to publication. There are so many who have written, called, volunteered services and done field testing that we may have inadvertently omitted acknowledging everyone. To you, we sincerely apologize. We especially acknowledge the guidance and initiative of Marjorie Gardner, who was one of the prime movers of this project until her untimely death in April, 1991. We dedicate *SourceBook* to the memory of this committed, dedicated teacher of teachers.

Mary Virginia Orna

James O. Schreck

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Editors of *SourceBook*

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The authors, editors and sponsoring agencies accept no moral or legal responsibility for injuries that may result from demonstrations or laboratory activities.

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

CHEMICAL PEDAGOGY

Topic Overview

> *Chemistry is terribly interesting! I never knew that things could change so much. I mean, you could have a rock one moment and a flower the next. They always show you that on cartoons. I was thinking, obviously, that's not possible, but when I got into chemistry, I saw that things comparable to that could happen! I was just completely amazed. I'd always wanted to know the chemical aspect of that.*
>
> Allison Garner

Interviews with students during videotaping of exemplary high school chemistry teachers in their classrooms last year, revealed diverse perspectives on learning and teaching of high school chemistry.

Many students, such as Allison Garner, thought that chemistry was great. Other students had reservations and appeared to be enrolled in the class to "please their parents," "get into college" or "become a doctor."

There are many reasons why students enroll in chemistry and why it is taught at the high school level. Sometimes these reasons are at odds with one another. Sometimes the students only have extrinsic motives to succeed. Isn't the dream of every high school chemistry teacher to have a classroom filled with students similar in attitude to Allison who was studying chemistry because she wanted to "know the chemical aspect" of the changes in the world surrounding her?

High school teachers give many reasons for teaching chemistry. These include teaching critical thinking skills such as problem solving, helping students acquire an understanding of the nature of scientific inquiry, and enabling students to appreciate and enjoy the world around them. In reality, however, many of these most worthy goals go unaccomplished because of external pressures to cover the book or to have students score well on examinations. Would it not be better to partially achieve the aforementioned goals at the expense of "covering the book" in order to have students emerge from the introductory chemistry course as excited about chemistry as was Allison?

Teaching is a very complex human activity. Human beings are very complex organisms. What motivates one student to learn chemistry may not motivate another student. For example, research shows that males and females have different learning preferences. If the purpose of teaching high school chemistry is to help students learn chemistry content, improve their critical thinking skills, appreciate the world about them, and develop or maintain a positive attitude toward science, steps must be taken to assure that the strategies teachers use do this for every student enrolled in chemistry courses. If students do not think that chemistry is fun, and if students are not intrinsically motivated to study, they will be less likely to succeed. This negative attitude not only reduces the likelihood of their pursuing a scientific career, but may be transmitted to their children and to future generations.

This chapter is about learning and teaching high school chemistry. A brief examination of what both research and students say about these topics sets the stage for the theoretical underpinnings on how students learn. Specific instructional strategies that teachers have found effective in motivating their students, will be presented for improving classroom instruction, enhancing problem solving, and facilitating laboratory learning.

THE PRESENT REALITY

Results of a recent international study on science achievement indicate that students in the United States score lower in chemistry than students in most other industrialized countries. Does this mean that students understand less chemistry? Just how much chemistry do students understand when they enter the chemistry class and how much do they understand when they emerge? How do students feel about chemistry? How do they characterize good teaching or effective teachers? Before discussing teaching strategies, it is important to know what students are learning in the chemistry classroom, and students' perceptions of good teaching.

RESEARCH PERSPECTIVES ON CONCEPT LEARNING AND PROBLEM SOLVING

In the past 15 years, numerous studies on concept learning and problem solving have indicated that students enrolled in science classes understand few chemistry concepts and solve chemistry problems using algorithmic methods. These deficiencies are not apparent from students' scores on standardized tests which frequently ask questions requiring factual knowledge or solutions to problems that can be solved algorithmically with little understanding. It is when students are interviewed and asked to explain their answers that their lack of conceptual understanding becomes apparent. For example, Yarroch (*1*) interviewed high school chemistry students who were able to balance simple equations correctly by inspection. He found that 50% of the students who balanced equations correctly could not represent the coefficients and subscripts at the particle level. These students represented 2 NH_3 as 000000 rather than as 000 000, or used some other incorrect representation. This is corroborated in a study using interviews by Lythcott (*2*) and in an analysis of carefully constructed multiple-choice test items by Friedel and Maloney (*3*).

Studies on children's conceptions of matter abound. A classic study by Osborne and Cosgrove (*4*) examined 12 to 17 year-old New Zealand students' conceptions on what happens to the chemical species of water as it boils, evaporates, and condenses. They found that while students' naive conceptions generally decreased with age, 40% of the 17 year-old students thought that the bubbles in boiling water were made of hydrogen and oxygen, 30% thought that water changed to oxygen and hydrogen when it evaporated, and 35% thought that hydrogen and oxygen combined to form condensation on the outside of a cold jar. In fact, some of the 17 year-olds even thought that the water seeped through the glass to form the condensation. Although it is difficult to believe that students in the United States could be so naive, a student in a chemistry class of a teacher who was being observed for the *SourceView* project asked out of the blue, "You know the water that appears on the outside of a cold glass, I have always wondered how the water got through the glass. Could you tell me?"

Conceptions that students have about relatively simple chemical concepts persist even after students have taken numerous chemistry courses. Bodner (*5*) reports that chemistry graduate students still have naive conceptions after an undergraduate major in chemistry. He found that almost 20% of the graduate students thought that the bubbles in boiling water consisted of air or oxygen and that 5% indicated that they contained a mixture of hydrogen and oxygen gases. He also reports on other misconceptions held by these students on such topics as temperature and heat, and rusting of iron and weight. Several examples of students' lack of expected conceptual knowledge are shown in the *SourceView* videotapes. In Episode Four, Ms. Briner asks students if it is possible to have a concentrated, weak acid? The question, which was asked after students had studied molarity and after a brief explanation of acid strength had been given, is answered "no" by a number of students. Another example is shown in Episode Five by Ms. McKibbin who asks her class what gas is being given off after she drops dry ice in cylinders of water containing indicators. Students are unsure of the answer and some indicate that it might be hydrogen, oxygen, or even ammonia.

Research on problem solving also indicates that students solve problems without understanding much chemistry. Greenbowe (*6*) concluded that undergraduate college students solving problems aloud frequently used an algorithmic approach, and they were unable to solve problems requiring unique solutions. Gabel, Sherwood, and Enochs (*7*) found not only that students could not solve transfer type problems, but students were unable to answer qualitative questions about them. For example, more students could solve molarity problems correctly using the formula, M = moles/liter, than could describe how to make up two liters of a 1.0 M solution of sodium chloride in the laboratory. In another study on problem solving at the college level, Bunce and Gabel (*8*) interviewed students on how they went about solving chemistry problems. When asked whether they thought about the physical phenomena involved, students indicated that this was not necessary. One just found the formula or pattern that fit, that is, the "Rolodex approach" to problem solving. Additional information on these studies and many other studies on conceptual understanding and problem solving have been reviewed by Gabel (*9*), Krajcik (*10*) and Gabel and Bunce (*11*).

STUDENT PERSPECTIVES ON CHEMISTRY TEACHING AND LEARNING

Chemistry teaching involves much more than lecturing to students who dutifully take notes and answer questions. It involves more than the performance of certain teaching strategies that will be described later in this chapter. Effective teaching is an art that depends not only on the depth of the content knowledge of the teacher, his or her ability to use appropriate strategies that motivate the students to learn, but also on human characteristics such as the personality of the teacher, and their respect for students.

As part of the *SourceView* project, students were interviewed on a variety of topics. Some of the questions asked focused on chemistry teaching; others asked students what they liked best about chemistry. Sixty-three students were asked what made an effective or exemplary teacher. The most common response (16 students) was the teacher had a sense of humor or made the class fun. The next most common response was that the teacher made the subject matter clear (10 students). This was followed by the comment that effective teachers made the class feel relaxed and students not afraid to answer (9 students). Interestingly, only four students mentioned that understanding the subject matter was an important characteristic of a good teacher. To sort the data, student comments were categorized into two groups according to whether they were related to the subject matter knowledge and the teaching methodology (such as using demonstrations and making comparisons), or according to whether they related to relationships between the teacher and students or personality characteristics (such as understanding, open-mindedness and enthusiasm). Students overwhelmingly cited the latter as characteristics of effective teachers. Of 146 comments made by students, 37 pertained to the content and methods whereas 109 were related to relationships and personality.

Three high school principals were also asked what they thought were important characteristics of effective teachers. It is interesting that two of three principals said that love, respect, and concern for students was the number one characteristic of an exemplary teacher. Two of three included a strong content base as another desirable characteristic, but this was ranked second and fourth. These data corroborate other research on teaching that excellent teachers not only know the subject matter well, but also have good rapport with students.

Certain aspects of chemistry are more appealing to high school students. When asked about what they liked best about chemistry, about "what turned them on," 45 of 64 students interviewed mentioned laboratory experiments. Many explained that it was exciting or interesting to actually see things happening. Twenty-five of this group said they liked chemistry because it was related to everyday life and helped them understand things around them. Other replies repeated by three or more students were problem solving (5 replies) and that chemistry was challenging (3 replies). On the other hand, when these same students were asked what they disliked about chemistry, the responses were more varied. The most disliked part of the course was problem solving (13 replies). This was followed by memorization (12 replies), equations (7 replies), lectures (6 replies), hard work (5 replies), formulas (4 replies), laboratory experiments (4 replies), and not understanding small things such as atoms, protons, quarks, *etc.* (3 replies). Interestingly, six students said there was nothing they disliked about chemistry.

Although this was not a random sample of all students enrolled in the videotaped chemistry classes, but consisted of students selected by their teachers for interviews, and thus might represent more articulate students, it is interesting that the laboratory and relating chemistry to everyday life was mentioned so frequently as reasons why students liked chemistry. This also makes a statement about what was occurring in the classrooms that were videotaped. Teachers in those classrooms were selected to participate in the project because they made few conceptual errors and had good rapport with students. They were considered exemplary teachers and most had also received or were nominated for teaching awards. Apparently in teaching, they made chemistry exciting by taking time relating chemistry to everyday life and having students perform laboratory experiments. Because of the importance of the laboratory, Part Three of the *SourceView* videotapes focuses on Laboratory Instruction. Numerous examples throughout all episodes relate chemistry to everyday life. Episode Eight, showing students in a role-playing simulation of a town meeting, is particularly effective in engaging students in a real world application.

VIEWS ON LEARNING

Over the years my chemistry teaching has changed. In the beginning, I felt the coverage of all topics in the book was most important, whether the students truly understood them or not. But now, I firmly believe that if I cover just a few concepts, and they have a genuine understanding of those concepts, that provides the foundation for the further advancement of their knowledge.

Karen McKibbin

Unfortunately, all chemistry teachers have not come to the same realization about how students learn chemistry as Ms. McKibbin, one exemplary teacher featured on the *SourceView* videotapes. Many teachers believe that if telling is teaching, then listening is learning. They present chemistry concepts as if students' minds are blank tablets. Lectures are delivered at a fast pace, and little time is allotted for student interaction, determining preconceptions students have about concepts, and engaging students in laboratory and other activities. Teachers using this approach should not be surprised that students neither learn nor enjoy chemistry. Today's views of how students learn differ from those that behaviorists proposed in the mid-sixties and that have influenced how teachers were encouraged to teach for the past 25 years.

An Information Processing View

Cognitive psychologists theorize that learning takes place in several steps. In general, individuals use their senses for the input of knowledge. Some of this knowledge enters the short term or working memory and eventually is stored in long term memory until it is needed for further use.

Much knowledge that one senses, however, is never remembered. For example, a teacher does not remember what each student is wearing even moments after looking at the class. Few teachers can remember all the students' names on the first day of class even after making an effort to do so! Certain information obtained through the senses is filtered out and other information filters through to short term or working memory. If a student is wearing an outlandish sweatshirt, or if a student has a nonconventional name, these generally aid the teacher in remembering what the student was wearing or his or her name momentarily.

Information that gets into short term memory remains there a very short period of time—about 15 seconds. In addition, many small pieces of information enter the short term memory at one time. How many people can remember their nine digit zip code the first time they hear it or read it, much less a chemical formula? Once information has entered into short term memory, it will either disappear or enter long term memory. Since one of the goals of education is to facilitate learning, it is important to understand ways information entering short term memory is transferred into long term memory. One way is through repetition. If a person repeats a zip code enough , or writes the formula for a chemical compound repeatedly, eventually it is remembered. Another way information is transferred is through association. If an individual can relate new information to be remembered with something that is already in long term memory, then it is likely to be stored and remembered. Information that is stored in long term memory is thought to be linked into networks of different complexity. It is hypothesized that one difference between expert problem solvers and novice problem solvers is that the experts have complex networks of concepts they draw upon to solve problems. Novices, on the other hand, have limited networks with little branching. Perhaps they can retrieve a memorized formula to solve a routine problem, but since the linkages to other concepts are not present, they lack the capability of solving problems requiring additional steps, or ones never encountered. From the information processing point of view, learning is enhanced by increasing organized networks of concepts that exist in long term memory.

The Constructivist View

From the constructivist's viewpoint, individuals create or construct their knowledge. This construction of knowledge begins at birth (or before) and continues throughout life. Hence children enter school with a conceptual understanding of phenomena that they have acquired by observing the world around them and trying to make sense of their observations. Usually this knowledge, called "spontaneous knowledge" or "personal knowledge" by West and Pines (*12*), differs from that of a scientist. For example, children (and some adults also) are surprised to observe that the temperature of a piece of velvet and a piece of metal in an air-conditioned room is the same. Since the metal feels colder, the senses trick one into believing something that is incongruent with what Pines and West refer to as "scientific knowledge" or "school knowledge." Personal knowledge is also created in other ways. For example, parents continually telling children to wear their "warm hats" may be the reason why most young children believe that the temperature inside a hat (when not being worn) is greater than the temperature outside a hat.

Unfortunately, what frequently happens in chemistry classrooms is that students do not make connections between personal knowledge and school knowledge. Moreover, if chemistry is taught to "cover the book," students do not even make connections between various aspects of school knowledge, and have little opportunity to relate it to personal knowledge. (A salient example of this has been reported for first grade children by Brenner (*13*). In a study of children in a middle-lower class neighborhood, five year old children were used to making change using quarters and dollars to purchase confections at the the corner grocery store. In first grade they were introduced to working with pennies. They saw little connection between the two systems of money. Five cents would only buy the lemonade cup, not the contents! What grocer would have the patience to allow children to count out 50 pennies for their purchase? The children developed two systems for dealing with money, a real world system, and a school system.)

One reason why students find chemistry challenging is because chemistry is a complex subject. In an effort to describe and make predictions about the world of matter, chemists explain what occurs on the macroscopic level using atoms and molecules. Moreover, they represent these atoms and molecules using symbols and formulas. To add to the complexity is the fact that the atoms and molecules are so small and exist in such large numbers that they are difficult for students to visualize because there are no meaningful examples in everyday life. This makes understanding chemistry even more difficult because of the necessity of using mathematics to describe and relate atoms and molecules to measurable quantities of masses and volumes at the macroscopic level. It is no wonder that high school students have difficulty in making the necessary connections among the many facets of chemistry!

However, chemistry teaching frequently does little to remedy the situation, and actually may add to it. As Johnstone (*14*) has indicated, most chemistry instruction takes place at the symbolic level, and little time is spent in helping students relate the symbols to the chemical phenomena and to the microscopic level.

This focus on almost an exclusive symbolic representation of matter in introductory science courses may be the cause of students' lack of understanding of basic chemical concepts as was illustrated in the studies cited earlier. A continual exposure in their high school and college courses to instruction on the symbolic and mathematical levels of chemistry will not lead students to make the connections between the science that they learn in the classroom and the world around them. How can we expect students to pursue a career in chemistry when the study of chemistry does not lead to understanding? However, if teachers are aware that this is the situation, something can be done to improve instruction. Gabel, Briner and Haines (*15*) have shown that when the particle nature of matter is emphasized in classroom instruction at the high school level, there is some evidence that student understanding increases on all three levels, even though the achievement levels were not as high as desirable. This emphasis on using particles to ask students to explain physical phenomena has been called using "pictures in the mind." In the *SourceView* videotapes, Mr. Haines uses "pictures in the mind" in Episodes Four and Sixteen when he shows students how to solve problems. In the episodes, Mr. Haines relates the physical phenomena with "pictures in the mind" with the symbolic and mathematical representations, thus helping students integrate the three ways of viewing matter.

Integrating the three ways of presenting chemistry is sometimes referred to as the SAP approach. Just as sap is an indication that a tree is alive and expected to grow in the future, only those students who understand chemistry on the three levels (Symbolic, Application or macroscopic, and Particle) have the conceptual base that is needed to be creative and to pursue a career in chemistry in their future.

A Conceptual Change Model

The goal of every chemistry teacher is to facilitate the understanding of chemistry concepts. Although some positive improvements may occur in students' conceptual understanding by teaching chemistry on the macroscopic, microscopic, and symbolic levels simultaneously, or by rearranging the content so that more concrete concepts are introduced before more abstract ones are presented as in the National Science Teachers Association's Scope, Sequence & Coordination Project (*16*), or by using an interdisciplinary approach such as that advocated by the American Association for the Advancement of Science's 2061 Project (*17*), the change will not be dramatic. Substantial change in students' understanding of science will only occur if teaching methodology and content are appropriate for students.

The use of the conceptual change model proposed by Krajcik (*10*), a modification of one proposed by Driver and Oldham (*18*), should lead to a higher level conceptual understanding of chemistry concepts than students are now experiencing in schools. The model consists of four steps that should be considered in teaching sequence. These include: (a) students describe their understanding, (b) students restructure their understanding, (c) students apply new understanding, and (d) students compare new understanding with previous understanding. The restructuring in step (b) includes student and teacher exchange to clarify understandings, student exposure to conflict situations through discrepant events, and student construction of new understanding.

Some teachers are unaware that they are already using parts of this conceptual change model. A sequence in the *SourceView* videotapes illustrates this process quite well. In Episode Four, Ms. Briner's major objective for the day is to lead students to understand the difference between strong and weak acids. Students have already studied concentration in the chapter on solutions earlier in the semester. Ms. Briner begins with a discussion of the previous day's laboratory exercise in which students serially diluted a strong and a weak acid and noted the effect on pH using indicators. After students conclude that there are differences in colors and the pH's of acids in different wells, one might think they understood that the concentration of hydrogen ion is less in the weak acid than in the strong acid when similar dilutions are made. However, there is evidence that this is not the case. Even after Ms. Briner provides a short explanation of differences between a strong acid and a weak acid in terms of percent ionization, when students are asked if it is possible to have a concentrated, weak acid, many students think that it is impossible. Even the spokesperson for the students who thought it was possible convinces himself that he was wrong! The question and discussion that ensues provides the cognitive conflict for the students in addition to making the teacher aware of the students' thinking. The activity that follows provides students with an opportunity to make appropriate distinctions between strength and concentration and to apply their understanding to a new situation. Ms. Briner deliberately has structured her lesson to make these distinctions because from previous experience she knows that students use these words interchangeably in everyday life and are likely to have a personal view of them not congruent with the scientific view. She knows that students will not understand this difference by just telling, and that they must confront their personal views on multiple occasions in a variety of contexts for true understanding. Otherwise students may simply memorize the definitions of "strong acid" and "weak acid," and be unable to apply them in everyday life.

CLASSROOM CLIMATE

> *An outstanding teacher is someone who makes the class feel relaxed, and makes the class participate. . . . He'll create an atmosphere where you're not afraid to answer, even if you are not quite sure. He is not going to scold you for being wrong. He's not going to embarrass you in front of everyone else.*
>
> Andy Hapskind

Classroom climate is extremely important for growth in understanding. Andy Hapskind and other students that were interviewed for the *SourceView* videotapes felt that way. Students are very reluctant to reveal their personal knowledge in a tense classroom situation. The personal qualities that students mentioned in exemplary teachers creating a positive learning environment were: sense of humor, friendly, fair, patient, kind, understanding, open-minded, enthusiastic, adaptable, calm, and love of students.

During the interviews, several students indicated that they learned very well in small group settings where they felt freer to make mistakes, and were able to listen to students' explanations of concepts. Research by Johnson, Johnson, Holubec, & Roy (*19*) and Slavin (*20*) has shown there are positive effects in achievement, retention, and critical thinking skills when students work in cooperative groups. As pointed out by Champagne and Bunce (*21*), who base their statement on a tenet of cognitive psychology proposed by Ausubel (*22*), peer explanations may be more effective than teacher-student interactions because "the distance between two students' understanding is far less than the distance between a student's understanding and a teacher's, hence communication of ideas is facilitated." With these and other views of student learning in mind, teaching strategies that engage students in learning are discussed.

VIEWS ON TEACHING

> *When I first started teaching, I taught like a college professor. I stood at the front and I lectured, and I do very little of that now. When I'm standing in front I'm questioning. I'm discussing with my students, and I don't stand in just the front all the time any more. I put them in groups. I make them the focus of attention. I put them in lab. I wander around. I talk with them. My teaching style is totally different now . . . seven years later.*
>
> Denise Briner

Ms. Briner, an exemplary *SourceView* teacher, engages students in learning whether the setting is the classroom, the lecture hall or the laboratory. Effective teaching strategies engage students in learning about the elements, solving stoichiometry problems or discussing a laboratory experiment. The discussion of effective learning strategies has been divided into three major components to correspond to Parts One, Two, and Three of the *SourceView* videotapes. These are: Classroom Instruction, Problem Solving, and Laboratory Instruction. Division of the teaching strategies into these three categories is somewhat arbitrary because some skills used for one of the components are also applicable in others, but by following this schema, the reader can make ready reference to the videotapes.

EFFECTIVE CLASSROOM STRATEGIES

As indicated in the section entitled "Views on Learning," students will learn best when there is good rapport between teacher and students. However, even when a favorable classroom climate exists, only a minority of students are able to make sense of the science that is presented in terms of their own prior conceptions when conventional instruction such as lectures is utilized. As pointed out earlier, most students will memorize definitions in lieu of understanding concepts. Classroom strategies presented in this section can be used effectively to promote conceptual change. They can also be used ineffectively.

Common to all strategies is the appropriate pacing of instruction. If students are not allowed enough time for processing information, either when asking a question or participating in a demonstration, no matter what the content or pedagogy, learning will not take place. Effective teachers build this "thinking time" into their instruction by using such techniques as asking students to close their eyes and to think about a particular situation, or pausing sufficiently after asking a question before calling on students for a response. If students are scheduled for large group lectures, the rapid pace of delivery can be modified by stopping for two minutes after every 12 or so minutes of presentation. Rowe (*23*) found that if students make use of this time by comparing notes with their near neighbors, and clarifying the material just presented with one another rather than spending the time asking the instructor questions, chemistry achievement improves. A healthy classroom climate and sufficient learning time are preconditions for effective implementation of teaching strategies discussed below.

Questioning

Probably the most commonly used teaching strategy at the high school level is asking students questions. When the technique is used effectively, it can aid students in concept acquisition, as well as provide feedback to the teacher about student comprehension. Four aspects about questioning need to be considered. These are: (a) What kinds of questions should be asked? (b) Why are questions asked? (c) How should questions be asked? and (d) How does the teacher respond to students' answers to questions?

Several schemes have been developed on question types appropriate for the classroom. For example, Cunningham (*24*) suggests that both narrow and broad questions are appropriate for classroom use, but serve different purposes. Narrow questions provide feedback to the teacher about what an individual student understands, whereas broad questions add the dimension of requiring some original thought on the part of the student. For example, Ms. Herron (in Episode Seventeen) uses a convergent question when asking what acid is present in vinegar. In Episode Ten, when asking students to tell how the objects on their chairs relate to acids and bases, the question is broad because there are many possible answers, and students may come up with a variety of responses.

A more complex schema that has been in use for about 30 years based on Bloom's (*25*) *Taxonomy of Educational Objectives* was proposed by Sanders (*26*) who classified questions according to seven levels: memory, translation, interpretation, application, analysis, synthesis, and evaluation. No matter what schema is used to classify questions, what is most important is that unless higher level questions are asked, very little thinking is occurring on the part of students. Students reply to low level questions with one or two words, perhaps a phrase. An observer in a classroom can easily tell the level of thinking occurring by simply observing the length of student replies to questions.

It is not only the level of the question asked, however, but also the content of the question that is important. If a teacher asks a question about a discrepant event, or about something that is counter-intuitive, conceptual change is more likely to occur. For example, Ms. Briner's question in *SourceView* Episode Four, "Is it possible to have a concentrated, weak acid?" begins a discussion that lasts about ten minutes in which all students are engaged in learning. Ms. McKibbin's question in *SourceView* Episode Five about what gas is being given off when dry ice is dropped into water produces a variety of responses. However, when Ms. McKibbin helps students realize that carbon dioxide is a clear, colorless gas, and the gas that they see is not clear, the questions she asks helps students resolve their cognitive conflict!

Teachers ask questions for a variety of reasons. As was just illustrated, one reason for asking questions is to get students to think. Questions asked before a demonstration will help students make predictions about the outcomes. After a demonstration, students can be questioned about inferences they can make. In both cases, students' replies provide the teacher with important information on their conceptual understanding and whether students' personal knowledge is congruent with scientific knowledge. Since each question is generally answered by only one student in the class, surveying students for answers to a given question as shown by Ms. Briner and Ms. McKibbin in the aforementioned *SourceView* episodes, indicates to the teacher the extensiveness of lack of congruence between personal and scientific knowledge among class members.

The methods that teachers use in asking classroom questions is also important. If a teacher calls on the student by name before asking the question, it is likely that other students in the class may not listen carefully to the question being proposed because they know they will not be required to answer. A more effective technique that gets all students involved in thinking is to ask the question, wait 3-5 seconds, and then call on a student by name to answer the question. Rowe (*27*) found that by including this "wait-time," students gave much longer responses than when questions were asked in rapid succession.

After a student has given a response, how should the teacher respond? Although individual teachers may have their own techniques for continuing the dialogue, one effective way to keep students involved in thinking is not tell them whether their response was correct. If the teacher then calls on other students for additional responses to the same question, comparisons can be made and class consensus can be reached. Obviously, only higher level questions can be used for this kind of discussion to ensue.

Another way to continue the class discussion is to reply to a student's answer with another question. Questions like, "Why do you think so?" or "How did you get that?" will encourage students to expand their thinking.

Using Analogies and Models

The use of analogies and models in teaching chemistry is quite common. An analogy can be thought of as the resemblance of otherwise dissimilar things, whereas a model is a physical entity usually built to scale to resemble something else. The use of analogies helps students construct meaningful knowledge by linking an unfamiliar concept with a familiar one. The greater resemblance in features of the two concepts, the more powerful the analogy. The effective use of analogies assumes that students are familiar with the analog. In a study of problem solving using analogies, Friedel *et al.* (*28*) found that a large number of students enrolled in a beginning chemistry course were unable to make connections between problems using familiar objects such as pieces of fruit, and problems involving atoms and molecules. Hence instruction using everyday objects was unsuccessful.

On the other hand, if students become involved in making up an analogy or if they participate actively in an analogous event, they are more likely to determine similarities and dissimilarities of the analog and that which it represents. An effective use of having students represent particles in a water solution is shown by Mr. Haines in *SourceView* Episode Six where students (particles) move about in different sections of the classroom (volumes of water) creating different concentrations.

Care must be taken to show the limitations of a given analogy. If students perceive that the analog and concept are identical in all aspects, they may inadvertently develop a misconception about the unfamiliar concept by attributing something to it that it does not possess. The same possibility occurs in using models. One reason why students may think that an individual atom of carbon is black may be due to the black color of the ball-and-stick model for the carbon atom! If the inventor of the ball-and-stick models had painted the model with four bonds green, this might have created cognitive conflict for students showing that the color and other properties of a large sample of atoms is not identical with the properties of a single atom!

The use of models aids students in understanding chemistry because they represent atoms and molecules too small to be envisioned by most students in a concrete way. Ideally, three-dimensional models represent molecules better than two-dimensional models. Irrespective of the type of model used to represent molecules, it is important students see the connection of the model to the microscopic structure of the compound, to the symbols and formulas which chemists use to represent atoms and molecules, and to physical phenomena as well. Effective uses of models are shown in *SourceView* Episode Three by Mr. Lumbley who uses them in a historical context, and in Episode Fourteen by Mr. Haines who depicts solution concentration using magnets in beakers that have been drawn on the chalk board.

Concept Mapping

In the past several years, concept mapping, as advocated by Novak (*29*), has become a useful strategy commonly used by teachers. A concept map consists of a hierarchical arrangement of propositions, that is, concepts connected by a verb that show the relationship between the concepts. For example, a part of a concept map on acids and bases might contain the proposition, "acids have a sour taste." In common use, the linking verb is frequently replaced by a line and the assumption is made that the linkage is understood. Figure 1 shows a typical concept map for acids and bases. Others on the same topic could be equally correct.

Studies on expert and novice problem solving behavior by Larkin (*30*) and Chi, Feltovich, & Glaser (*31*) indicate that experts appear to have complex, elaborated conceptual schemes in long-term memory that they draw upon in solving problems. These appear to be lacking in novices. Since the construction of a concept map on a given topic helps students organize concepts and the relationships among them in a physical way on paper, this may facilitate the construction of a similar conceptual map in students' long-term memory, and hence be a valuable tool for promoting conceptual understanding.

Concept maps can be used in many ways and care must be taken that students do not resort to memorizing maps. They can be constructed and used by the teacher in planning a unit. Students can be asked to construct maps from content in a short section of their textbook. A very creative way of using maps is illustrated by Ms. Herron in Episode Ten in the *SourceView* videotapes. At the end of the unit, students are asked to construct a map from concepts linking real world acid and base phenomena. The creation of the map not only serves as a review for students in helping them construct an overview of the unit, but aids Ms. Herron in identifying those concepts about which students need additional help (*41*).

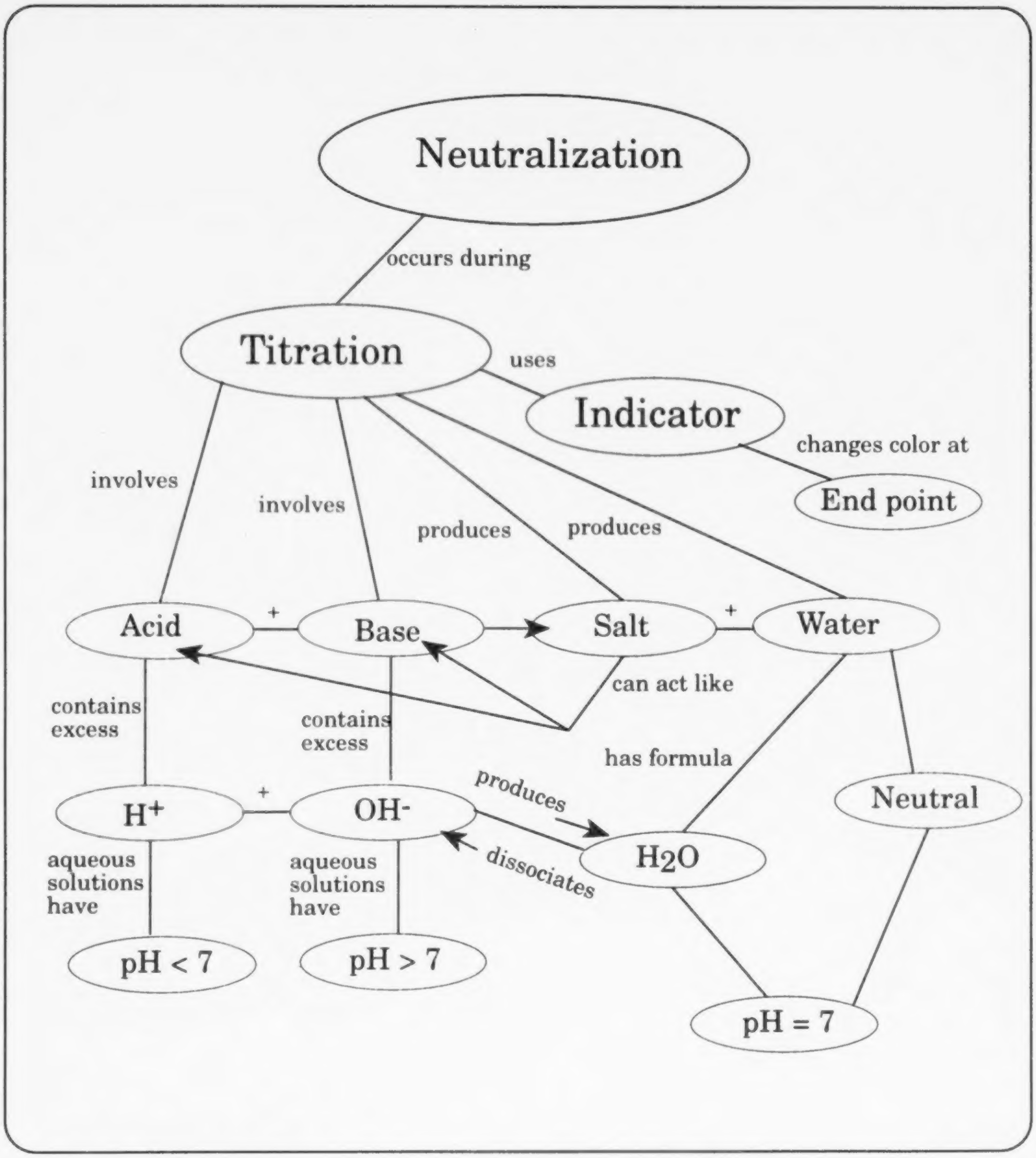
Neutralization
occurs during
Titration
uses
Indicator
changes color at
End point
involves
involves
produces
produces
Acid
+
Base
Salt
+
Water
contains excess
contains excess
can act like
has formula
H+
+
OH-
produces
dissociates
H2O
Neutral
aqueous solutions have
aqueous solutions have
pH < 7
pH > 7
pH = 7

Figure 1. Concept Map for Neutralization

Collaborative Learning

> *Often times I'm scared to ask a question to the teacher if I'm not quite sure of everything, but when you're working with your own peers, you feel better about asking them a question, because they often times know the answer, and they can tell you better. . . .*
>
> Philip Larson

Many students feel the same way as Philip Larson. When students work with each other, they feel more at ease and freer to ask questions about concepts needing clarification. This is in addition to the benefit discussed earlier of the student receiving an answer that is more understandable because the concept development and language used in the explanation of one student is more similar to another student's than to that of the teacher. Assimilation of new ideas is contingent on how closely the new ideas are related to present knowledge.

Chemistry laboratory usually provides students with many opportunities to work together. In the *SourceView* survey, working in the laboratory was what students liked best about chemistry. However, there are many other ways that chemistry instruction can be structured to allow for student collaboration. Even in a large group lecture setting, students profit if the instructor stops every 12 minutes or so and allows students to discuss the content that has been presented with one another.

A small group setting of 24-30 students provides more opportunity for students to work in small groups. The goals to be accomplished will determine the type of collaboration that is appropriate. If the objective of the lesson is to have students become proficient in a skill, such as learning to balance equations, students can work with their neighbor in appropriately designed exercises. Ms. Dillon's use of small groups in determining salt character shown in *SourceView* Episode Seven illustrates this technique very well. The sequence of instruction is as follows: Mrs. Dillon presents and models the thinking needed for skill acquisition at the board, students work in small groups, Ms. Dillon circulates among students to encourage, help, and determine difficulties, and then, summarizes and clarifies difficulties at the board. The process is repeated several times throughout the period.

At other times, it may be beneficial to include small group work later in the class period for a longer period of time. In Ms. Briner's class (shown in *SourceView* Episode Four) small groups of students use powdered drink mix to help clarify the difference between "concentrated" and "strong." In Episode Fourteen, Mr. Haines' students use worksheets to solve concentration problems.

Another effective way for students to work cooperatively is in the use of the jig-saw technique. It is appropriate for student research on various aspects of a given topic. For example, an appropriate use in a unit on acids and bases would be in the study of acid rain. The technique consists of dividing the class into groups of three to five students. Each student in the group studies a different aspect of the topic. These topics are predetermined either by brainstorming, some preliminary research by the class, from the textbook, *etc*. Some appropriate topics for the study of acid rain if the class is divided into groups of five might be: causes, geographical occurrence, effect on living things, effect on statues and buildings, and solutions to the problem. The students do not research these topics alone, however. Students who have selected or who were assigned the same topic now meet as a group. For example, if there were 25 students in the class, and five groups of five students were formed, the student from each group responsible for "occurrence" would meet together. They would plan the approach they would take, and perhaps divide the topic further. They would continue to meet in this specialty group until they had researched the topic thoroughly, and shared their information with one another. Students would then return to their initial group where they would be responsible for teaching students what was learned about their specialty (*e.g.* occurrence) to other group members. In this way, each student is responsible for conveying information to his/her peers, and each student is responsible for learning about all aspects of the topic (*e.g.* acid rain).

Role-playing provides another opportunity for students to work together on various aspects of a topic. Students can work cooperatively in doing the necessary research that precedes their presentation, in planning and producing appropriate visuals, and in practicing and making their presentation. *SourceView* Episode Eight shows students working together effectively in several of Mr. Escudero's classes as they plan and present a role-playing scenario of a town meeting on a fish-kill.

Making Real World Applications

> *The thing I like best about chemistry class is that it helps me get a better understanding of the world around me, and I comprehend things I didn't know earlier or before I had taken the class.*
>
> Benji Lyn

After the laboratory experiments, students who were interviewed for *SourceView* most frequently mentioned "real world applications."It is easy to understand why this is so. There is a certain intellectual satisfaction when a person finally sees how concepts relate and everything falls into place. In terms of learning, discrete, independent networks in long term memory are being linked together to form a much larger, integrated network. This increases the students' ability to solve problems and results in a certain euphoria. Even if chemistry is taught in such a way as to link the atomic, phenomenal, and symbolic levels, unless new material is linked to the concepts already in long term memory, it will be compartmentalized into school or scientific knowledge *versus* personal knowledge. This, of course, results in students never making the connections between what is taught in school and the real world, and results in students who are unable to answer even the simplest questions about the world in which they live, as was illustrated by Bodner (*5*).

There are many ways to connect chemistry to the students' world, and almost every *SourceView* episode illustrates this. The role-playing of the town council meeting by Mr. Escudero's class in Episode Eight illustrates a real world application that takes three class periods. The laboratory experiment performed by Ms. Herron's class in Episode Seventeen and by Mr. Mill's class in Episode Twenty relates the value of chemistry in determining which vinegar is the best buy. Ms. Walsh in Episode Two and Mr. Lumbley in Episode Three relate the properties of acids and bases to household reactions and foods. Ms. Briner uses powdered drink mix to explain solution concentration in Episode Four and Ms. Goshorn relates acid rain to car emissions and industrial pollution in Episode Nine.

Doing Demonstrations

Demonstrations also provide opportunities for students to make connections to the real world. Demonstrations that use familiar materials that act in unexpected ways provide an element of surprise and cognitive conflict to make them effective. For example, a demonstration involving changes of colors of indicators such as blueberry juice or cabbage juice in acids and bases, as shown by Ms. Walsh's class in *SourceView* Episode Two, is more interesting to students than the changes with litmus or phenol red.

This is not to say that unfamiliar chemicals should never be used. Sometimes there is no effective alternative, and their use broadens students' familiarity with the material world. Even chemicals that the teacher might think are familiar may not be, and doing the demonstration gives the students an opportunity to understand common materials. When Ms. McKibbin in Episode Five planned the demonstration to show the effect of acids on indicators, she assumed that students knew that the gas coming from the cylinders was carbon dioxide with water vapor condensing on it. It was only by asking students about the phenomenon that she determined that this was not the case, and she seized the moment to revamp the demonstration. If she had not asked questions, and had merely told the students what was happening, very little learning would have taken place.

For an effective demonstration, presenting interesting phenomena is insufficient. The manner in which the presentation takes place determines effectiveness. Students should be involved in making observations, predictions, and inferences. Effective teachers sometimes "lead students on" by asking the correct questions without revealing the outcomes. Mr. Lumbley does this in Episode Three. Sometimes, it is appropriate to perform a "silent" demonstration in which students are asked to make observations while the teacher performs the demonstration silently. Mr. Haines' students in Episode Six are fascinated as he demonstrates the differences in reactivity of a concentrated and a dilute acid. In this instance, inferences about what happened occur after the demonstration has been completed.

Using Technology

Many technological tools are now available for teaching chemistry that hold promise for promoting conceptual change. These include microcomputer-based laboratory experiments (MBL), computer simulations, graphing packages, videodisks, and videotapes. A variety of these excellent instructional aids are listed in *SourceBook* modules and are shown being used by teachers in *SourceView* videotapes. The same can be said about the use of these tools in promoting conceptual change as was said about the use of demonstrations. If they are not used properly, learning will not occur. Two excellent examples occur in Episode One, a simulated lesson of poor teaching techniques, in which Ms. Walsh deliberately shows the incorrect use of two excellent instructional tools. She begins the acid-base unit with a short segment of the **World of Chemistry** videotape *The Proton in Chemistry* (*32*). Using this sequence at the beginning of the lesson could have provided an excellent unit introduction to stimulate student interaction about how acids relate to everyday life. Instead, however, Ms. Walsh uses it as a vocabulary-building exercise! Later in the period, she shows the **Project Seraphim** *Periodic Table* videodisk (*33*). Again, she deliberately misuses this excellent resource by telling students everything that they see, rather than calling on students to make the observations and offering explanations. Examples of excellent uses of a videotape and a videodisk are shown in other

SourceView episodes. Ms. Goshorn shows a segment of videotape Nine that provides input for the class discussion on the effects of acid rain. Ms. Herron uses the **American Chemical Society's** *Doing Chemistry* videodisk (*34*) to provide input on proper laboratory techniques for titration in Episode Seventeen. Appropriate student involvement is necessary for the successful implementation of technological tools.

Assessing Student Learning

Assessing student learning throughout instruction is a requirement of good teaching. If one accepts the fundamental premise that students construct their own meaning, it is important for teachers to assess students' existing knowledge in order to plan appropriate learning activities. Sometimes students will give nonverbal cues that indicate that they do not understand a concept. Other times the teacher will become aware of students' knowledge by circulating among study groups, or by asking appropriate questions. These methods have been discussed in previous sections of this chapter.

Assessment that occurs at the end of a study unit frequently takes the form of a written test. Unfortunately, tests that accompany many high school chemistry books use a short answer or multiple choice format. In addition, many of the questions that they contain only assess whether students can memorize the content well or are able to solve problems using algorithms. Just as standardized tests drive teachers to include certain chemistry concepts in the curriculum, the content and format of tests that students take drive them to learn chemistry in a rote and algorithmic manner. If the desired outcome of chemistry education is the conceptual understanding of chemistry, more open-ended tests that require explanations are needed. These kinds of tests which require more student thinking also require more administration time, and more time to correct. Alternative ways of assessing students' understanding are also necessary. These may include laboratory practicals, and the interviewing of students. When students understand that they are responsible for their own conceptual understanding, and that they will be assessed and rewarded on that understanding, conceptual understanding will be enhanced.

ENHANCING PROBLEM SOLVING

> *When people say they can't do the math in chemistry, what they mean is, they can't take a physical situation, read that situation, visualize it, and then write down some sensible, coherent, mathematical expression that can be solved to find what they're looking for. That's what they mean, and that's a challenge.*
>
> Frank Cardulla

In this brief statement, Mr. Cardulla says it all! Problem solving in chemistry involves more than getting answers. Meaningful problem solving, that is, problem solving that involves more than recognizing and using appropriate algorithms, requires students to follow a series of steps to arrive at a reasonable answer. To help students achieve this is one of the greatest challenges of chemistry teaching.

Unfortunately, as indicated in the section on Research Perspectives on Concept

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Learning and Problem Solving, most students solve problems in a very algorithmic manner, and do not relate the problem solution to any physical phenomenon. Because students are unable to make sense of the problems, it is no wonder that problem solving was mentioned most frequently by students interviewed for the *SourceView* videotapes as the thing they liked least about chemistry. The challenge for chemistry teachers is to make problem solving more meaningful so that it aids conceptual understanding rather than hinders it by adding to the memory load.

To do this, chemistry teachers need to encourage students to think rather than to rely on algorithms exclusively. Students want to know the formula! They would prefer to set up factors with units, cross out the units, and arrive at a numerical answer without relating the problem to its underlying concepts or to the real world. However, in allowing this to happen, students never realize the intellectual satisfaction of seeing how everything makes sense.

Integrating Johnstone's Levels (*14*) into Problem Solving

In the section on the Constructivist View, three levels of chemistry were discussed. These were: the phenomenological level, the atomic/molecular level, and the symbolic level. Frequently in an effort to "cover the content," problems are solved almost entirely on the symbolic level using both chemical and mathematical symbols. This is illustrated in *SourceView* Episode Eleven. In contrast, Episode Twelve shows a conceptual approach in which Mr. Cardulla begins with a demonstration using actual solutions. Before working with problems on the symbolic level, he relates concentration to something concrete such as raisins. The raisins serve as models for atoms, ions, or molecules. Thus the three ways of referring to matter are integrated into the lesson. Mr. Haines in Episode Fourteen also relates the symbolic to the particle level using magnets to represent the particles in a microscopic volume of solution that has been enlarged sufficiently to count the particles.

Another way of getting students to connect a phenomenon to its symbolic representation is by asking them to make up a problem for the class to solve as is shown in *SourceView* Episode Thirteen. It is not surprising that students may suggest that an appropriate problem would be to find the molarity of a solution made by dissolving 100 grams of salt in 10 milliliters of water! Including additional information in problems may help students think more of the physical reality. Creating problems that include variable names rather than units may also do this. For example, if a problem reads, "How many milliliters of 0.1 M hydrochloric acid . . . . " is changed to "What volume of 0.1 M hydrochloric acid . . . . ," students must connect the variable name "volume" with the units "milliliters," and perhaps make a better connection with reality.

Using A Systematic Approach

Research by Bunce *et al.* (*35*) has shown that if students use a systematic approach in solving problems they are usually more successful. Mr. Cardulla identifies the appropriate steps necessary for successful problem solving in the introductory quotation.

The first step is to "read that situation." This appears much easier than it actually is. Cognitive psychologists would call it "problem restructuring." It consists of decoding or translating the words given in the problem in a meaningful way. There are many barriers that may prevent the restructuring to occur in an expert-like manner. Students may be unfamiliar with the vocabulary, or the format may contain too many variables for the limited space available in a student's short-term memory. Students' problem solving ability generally improves when simpler words are substituted for more complex ones. Writing down what is given in a problem and what is sought generally extends students' short term memory.

Another step recommended by Mr. Cardulla is "visualizing it," that is connecting the words in the problem with the physical phenomena that are represented. In most cases, chemistry problems can be solved by simple arithmetic. If students understand the physical situation of the problem, they can compare it to a similar problem with which they are familiar, and solve it in the same way. This makes them less dependent on setting up problems using the factor-label method in a rote manner so that the units "cancel" out. Mr. Cardulla contrasts the two methods in Episodes Eleven and Twelve.

After students have read, and visualized the situation, they should be able to write down a mathematical expression and solve the problem. This is easier said than done, and it aids students greatly in this process and in the other steps given by Mr. Cardulla if the teacher or other students model their thinking by solving problems aloud. *SourceView* Episodes Eleven through Fourteen show teachers solving problems aloud and Episode Thirteen shows a student in action.

When students first begin solving chemistry problems, it is frequently helpful for them to use a guide that aids them in using a systematic approach. The one used by Mr. Haines in *SourceView* Episode Fourteen includes the following components: what is given, what is sought, other information that must be recalled, diagram showing phenomenon and particles, planning, mathematical solution, and check. When students use this guide they approach the problem in a systematic manner and simultaneously relate the phenomenological, atomic/molecular and symbolic levels of chemistry.

Collaborative Learning

As indicated in the section on Classroom Climate, students achieve more in group settings than when they work alone. This is particularly true for problem solving because students need to hear other students' explanations of how they solve problems. In addition, students feel freer to make mistakes in the group setting where they won't be censured by the teacher. Students solving problems collaboratively are shown in *SourceView* Episode Fourteen. Other suggestions for teaching problem solving that are based on research are found in Gabel (*9*), Woods (*36*), and Gabel and Bunce (*11*).

FACILITATING LABORATORY LEARNING

The thing I like best about chemistry are the labs because I think they get across the point on the chapters, and the concepts that they're trying to explain.

Margie Greene

Chemistry teachers have always seen value in having students perform experiments, and students find them the thing that they like best about chemistry. In *SourceView* interviews students frequently commented about how it was the laboratory exercises helped them relate chemistry to the real world. On the other hand, the laboratory component of the chemistry course is costly, and research has shown that performing experiments doesn't necessarily improve achievement.

Types of Laboratory Exercises

Laboratory exercises or experiments can be categorized in many ways. One way to classify them is according to the amount of chemicals that are used. In the past few years, more teachers are having students perform experiments using small-scale techniques. The advantages of using small quantities of chemicals include reducing cost, saving time needed to prepare large quantities of solutions, increased safety because the chemicals are generally more confined, and less pollution of the environment in their disposal. Students can usually observe the chemical reactions just as well using small quantities, and frequently students exert more care in using them because a small error can make a big difference in the results. In addition, the experiments generally proceed faster, and there is usually time to perform several trials in one class period. Small-scale techniques are shown in *SourceView* for the serial dilution of an acid and a base in Episodes Sixteen, Eighteen, and Nineteen.

Another way to classify experiments is according to their purpose. Sometimes an experiment may simply be an exercise to demonstrate a laboratory technique or to illustrate a given phenomenon that leads to concept development. The serial dilution shown in *SourceView* Part Three is this type of exercise. It can be used effectively to illustrate changes of indicator color according to concentration. It is quite different from the titration of generic and brand name vinegars which serves not only to review the concepts and psychomotor skills needed for understanding and performing the titration, but also to help students make value judgments about which vinegar is "better." *SourceView* Episodes Seventeen, Twenty, and Twenty-one show students conducting a titration experiment.

Both experiments are important in the chemistry curriculum, and both can be used to help students acquire inquiry skills. It is important that students learn to make multiple observations, realize that there are many correct answers, and determine if their observations and data are internally and externally consistent. These and other aspects of scientific inquiry are evident in *SourceView* Episodes Seventeen through Twenty-one. Episode Twenty-one is very effective in involving students in the scientific approach to what should be done when results go awry.

Timing of the Laboratory Experiment

One reason why students may not profit fully from a laboratory exercise is because it is placed in an inappropriate place in the curriculum. If the laboratory is not available to the class on a daily basis, this problem may be impossible to overcome.

The purpose for conducting the laboratory experiment will determine its placement in the daily schedule. Sometimes the experiment should occur before concept development has taken place. In these instances, the concepts flow from the experiment. For example, Ms. Dillon's students in Episode Sixteen are given only the directions for the serial dilution of the acid and base in the pre-laboratory because her post-laboratory discussion uses the information from the experiment to invent the concept "pH." In Episode Nineteen, which portrays a post-laboratory session, Mr. Escudero illustrates this. On the other hand, it would be foolish to begin a lesson on acids and bases with a titration of generic and brand name vinegars. It is more appropriately scheduled as an application toward the end of the unit after students know the meaning of titration as is shown in Episodes Seventeen and Twenty. Otherwise, there would be so much for students to learn (beyond the capacity of short-term memory) that they would merely go through the motions of performing the experiment by following directions.

A method that includes experiments that has been found to be effective in promoting conceptual change is the learning cycle approach. This approach was initially proposed for the teaching of elementary school science by Robert Karplus and was used in the *Science Curriculum Improvement Study* (*37*). Abraham and Renner (*38*) adapted it for use in chemistry instruction at the high school and college levels. Instruction is composed of three phases: exploration, invention, and application. In the exploration phase, students conduct laboratory exercises that help them explore the phenomenon under consideration. For example, beginning an acid/base unit by testing a variety of foods with indicators (as shown in Episode Two) or having students taste a variety of foods (as shown in Episode Three) might constitute the exploration phase. In the invention phase, the concept is deduced from the experiments performed during the exploration. In this case, students would invent an operational definition of an acid and a base. Ms. Walsh's students are shown doing this in Episode Two. During the final phase, students apply what they just learned to a new situation. In the titration experiment (Episodes Seventeen, Twenty, and Twenty-one), the titration of an acid with a standard base constitutes the exploration phase, a discussion of neutralization would occur in the invention phase, and the titration of vinegars or antacid tablets might constitute the application phase.

Pre-Laboratory Discussion

The purpose of the pre-laboratory session is to enable students to carry out the experiment in a meaningful way. It should always include a demonstration of new laboratory techniques, a reminder about safety precautions, and directions for the disposal of hazardous chemicals.

Additional information discussed will depend on the purpose of the experiment. In most instances, students will learn more if there is an element of surprise in the results. This tends to create cognitive conflict. Good examples of different types of pre-laboratory situations are shown in *SourceView* Episodes Sixteen and Seventeen. As mentioned previously, Episode Sixteen shows the pre-laboratory for the serial dilution of an acid and a base. Its focus is on establishing proper techniques. Episode Seventeen shows the pre-laboratory session for the vinegar titration. Its purpose is to demonstrate how chemistry is useful in making a value judgment about which vinegar is "better."

Conducting the Experiment

Peer and student-teacher interactions are important during the laboratory session. The teacher's role is to circulate among students asking them questions, keep them on task, and ensure that safety precautions are observed. In Episode Fifteen, Ms. McKibbin shows what not do to, whereas in Episode Eighteen, she demonstrates appropriate teacher behavior.

The chemistry laboratory provides an ideal setting for peer interaction, the importance of which was discussed previously. Cooperative learning, as proposed by Johnson and Johnson (*39*), is an effective method providing for peer interaction while simultaneously simulating how scientists might work cooperatively in real world situations. Instead of each student doing all parts of the experiment, students work in teams of four, and submit a joint laboratory report for which they all receive the same grade. Each member of the team has a specified role. Although these may differ slightly from one teacher to the next, four common roles that are generally assigned are: the experimenter, the materials manager, the foreman, and the recorder. The materials manager gathers materials at the beginning of the laboratory session and returns them at the end. The experimenter carries out the physical part of the experiment with the help of the foreman who keeps everyone on task. The recorder records the data and is responsible for the report. *SourceView* Episode Sixteen shows Ms. Dillon reviewing the roles with her class. Students are assigned to groups either randomly or according to ability so that mixed ability groups result. Roles rotate by week or by experiment in order to give each student a wide range of experiences.

Cooperative learning has proven to be a successful laboratory management technique. It simulates a real life situation, results in on-task discussion by students, helps students improve their social skills, and reduces the experiments and chemicals needed by one-fourth. Because the use of cooperative groups results in fewer laboratory reports, the teacher has time to evaluate them in more detail, and thus provide better feedback to students.

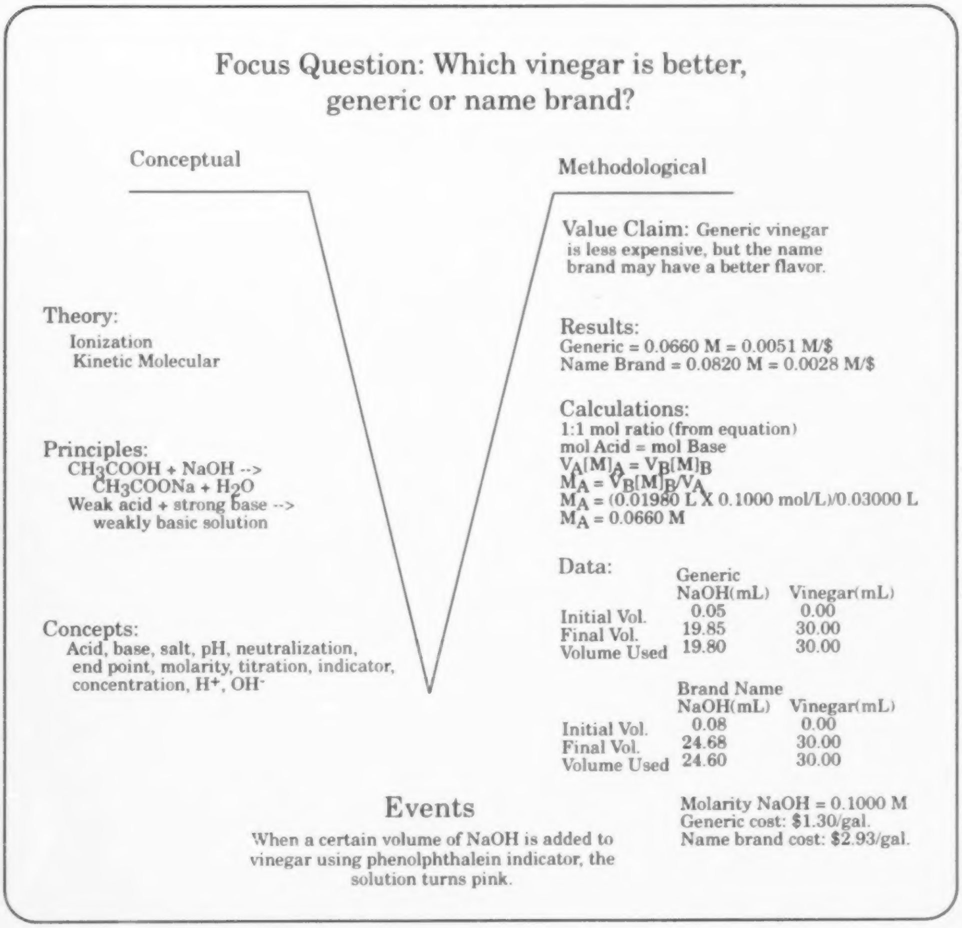
Post-Laboratory Discussion

The discussion that follows the experiment determines its real value as a learning experience. This can easily be seen by contrasting the post-laboratory discussion shown in *SourceView* Episode Fifteen with those shown in Episodes Nineteen, Twenty and Twenty-one. In Episode Fifteen, Ms. McKibbin simulates poor teaching behaviors by rushing through the post-laboratory discussion. Students probably would have learned as much without doing the experiment by just observing her results. This is in contrast to the discussion shown in Episode Nineteen in which students compare and contrast their findings.

The discussion of laboratory results provides the teacher with an opportunity to help students understand the nature of scientific inquiry. Such things as the need to repeat experiments numerous times, the necessity of examining the data for discrepancies, the meaning of significant figures, the necessity to play sleuth, and the need to repeat parts of the experiment with variations that may provide more information, cannot be shown or learned outside the context of a real experiment. These and other facets of scientific inquiry are illustrated by Mr. Cardulla and his class in *SourceView* Episode Twenty-one.

An effective discussion of experimental results also leads students to link what has been learned in the laboratory to what is being learned in the classroom. Unfortunately, many students do not make this connection. The use of the Vee Diagram as suggested by Gowin (*29*) to synthesize laboratory results is an effective tool for helping students make the link. A typical Vee Diagram, as shown in Figure 2, is divided into two halves. On the right side, is found the same components that would be included in a conventional laboratory report. On the left side are included the concepts, principles, and theories that link the laboratory experiment with prior learning.

The effective use of a Vee Diagram is portrayed in *SourceView* Episodes Seventeen and Twenty. In Episode Seventeen, Ms. Herron establishes the focus question with her class during the pre-laboratory session. Mr. Mills uses the Vee diagram with his class in Episode Twenty as the basis of the post-laboratory discussion and for the students' laboratory report.



Focus Question: Which vinegar is better, generic or name brand?

Theory:
Ionization
Kinetic Molecular

Principles:
$CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$
Weak acid + strong base --> weakly basic solution

Concepts:
Acid, base, salt, pH, neutralization, end point, molarity, titration, indicator, concentration, H^+, OH^-

Value Claim: Generic vinegar is less expensive, but the name brand may have a better flavor.

Results:
Generic = 0.0660 M = 0.0051 M/$
Name Brand = 0.0820 M = 0.0028 M/$

Calculations:
1:1 mol ratio (from equation)
mol Acid = mol Base
$V_A[M]_A = V_B[M]_B$
$M_A = V_B[M]_B/V_A$
$M_A = (0.01980 \text{ L} \times 0.1000 \text{ mol/L})/0.03000 \text{ L}$
$M_A = 0.0660 \text{ M}$

Data:

Generic	NaOH(mL)	Vinegar(mL)
Initial Vol.	0.05	0.00
Final Vol.	19.85	30.00
Volume Used	19.80	30.00

Brand Name	NaOH(mL)	Vinegar(mL)
Initial Vol.	0.08	0.00
Final Vol.	24.68	30.00
Volume Used	24.60	30.00

Molarity NaOH = 0.1000 M
Generic cost: $1.30/gal.
Name brand cost: $2.93/gal.

When a certain volume of NaOH is added to vinegar using phenolphthalein indicator, the solution turns pink.

Figure 2. Diagram for Comparison of Generic and Name Brand Vinegars Using Titration

CONCLUSION

Teaching and learning are very complex processes. This chapter has focused primarily on some of the processes and teaching strategies related to engaging students in learning chemistry during the time students interact with one another and the teacher. Less emphasis has been placed on other very important activities that occur before and after instruction. These include such things as planning, evaluating teacher effectiveness, and the professional development of teachers.

This chapter has attempted to consolidate and synthesize information on what constitutes exemplary teaching from three perspectives: those of chemistry students, those of chemistry teachers, and those of science educator researchers. We have seen that effective teachers engage students in learning by creating a classroom atmosphere in which students want to learn, and that it is important for teachers to establish a good rapport. Students highly value teachers who have a sense of humor, and make learning fun.

This chapter has not addressed all factors that characterize competent teaching. In a detailed and excellent review of the general research base of what are called preactive, interactive and postactive teaching tasks, Reynolds (*40*) elaborates on many other factors. These include such things as maximizing the time spent on worthwhile activities, making appropriate transitions between activities so the lesson maintains integrity, and establishing rules and routines to establish order that are perceived to be fair. For a more extensive review of these and other factors integral to effective teaching, the summary is highly recommended.

Reynolds (*40*) also indicates the importance of creating lessons that enable students to relate what they already know to new information. This requires teachers not only to know the subject matter, but to also know the students in order to tailor instruction to their needs. The domain of understanding for effective chemistry teaching is very comprehensive. It includes knowledge of the liberal arts, general principles of teaching and learning, and chemistry content. The integration of these three domains is necessary for expertise in the fourth domain that is sometimes referred to as "pedagogic content knowledge."

"Pedagogic content knowledge" refers to knowledge of specific ways of engaging students in acquiring a particular concept. In this chapter, an attempt has been made to describe some specific ways of effectively engaging students in learning. Some of these strategies will be more effective for certain concepts than others. It is only the skilled chemistry teacher who has a repertoire of strategies at his/her fingertips, that can make the judgment about the best method for presenting a particular concept to a given group of students.

In the *SourceView* videotapes, effective strategies to teach concepts pertaining to acids and bases to specific groups of students are presented by high school chemistry teachers. What works effectively in their situations may not work well in yours. All of this leads to the conclusion that becoming an expert teacher is a lifelong task of learning and practice. Even teachers who are recognized by their peers as excellent, are in agreement that it is important to continually expand their knowledge base in order to engage students in learning more effectively. Adopting and adapting the teaching strategies shown in the *SourceView* videotapes and discussed in this chapter should facilitate beginning, competent, and expert teachers in their life-long endeavor of helping students understand chemistry.

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A SourceBook Module

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

USER'S GUIDE

ChemSource Overview

ChemSource is an array of linked products consisting of two major components: *SourceBook* and *SourceView*.

SourceBook is a resource providing specific teaching tips and the best available instructional ideas and information gleaned from experienced, successful chemistry teachers. The core of *SourceBook* is thirty-six content modules featuring in each: Content-in-a-Nutshell, Laboratory Activities, Demonstrations, Common Misconceptions, Humor, Applications and Implications, Analogies, Reference resources, History, and Media. General materials on pedagogy, classroom management, safety, teaching chemistry to the physically disabled, and other useful resources and tabulations precede and follow the content modules.

SourceView consists of five-plus hours of videotapes of exemplary chemistry instruction at the high school level. Twelve teachers and their students are portrayed in twenty-one learning/teaching episodes in their actual classrooms. Sections centered around the content area, Acids and Bases, include Beginning, Presenting and Ending the Lesson, Problem Solving, and The Laboratory. The accompanying teacher's guide includes sections on Suggestions for Use, Teaching Skills Inventory, Overview and Bibliography, Reproducible Coding Sheet, Lesson Outlines and Transcripts, Episode Questions and Answers, Coding Sheet Answer Key, and a time coded matrix for indexing the lessons involving particular skills. The videotapes may be used individually by the novice teacher or a student teacher/cooperating teacher team as a resource for teaching techniques or by preservice chemistry teachers in university methods classes as a tool for developing teaching strategies.

- The primary purpose of *ChemSource* is to provide chemistry teachers with quality support to enhance their effectiveness in the classroom. *ChemSource* can be useful to beginning chemistry teachers who sometimes find the experience of preparing for teaching overwhelming. In many cases, these teachers have been called upon to cross over into chemistry from other disciplines, and while they may be experienced teachers in other areas, chemistry is a new field to them. In addition, many schools have only one science teacher, so the built-in chemistry mentors in larger institutions are not readily available in the one-science-teacher school. As a day-to-day resource and support strategy, *ChemSource* can overcome many of the problems facing the teachers described above. Compiled by widely recognized high school and college teachers, the *ChemSource* materials contain ideas, demonstrations, laboratory activities, teaching techniques and strategies that experienced teachers have found really work.

- The second purpose of *ChemSource* is to provide all teachers, both new and experienced, at the high school and introductory college level, with relevant, up-to-date material over a broad range of chemistry curriculum topics. Approximately one-third of the topics covered in the *SourceBook* component of *ChemSource* relate to applied chemistry or are enrichment topics which supplement the main curricular topics in the first year chemistry course.

- The third purpose of *ChemSource* is to provide a comprehensive resource that can be applied in numerous ways: in methods courses for the preservice teacher; in summer institutes or teacher workshops as an in-service resource; as an idea bank for enrichment and research projects for TAG (Talented and Gifted) programs; as a supplement for Honors chemistry programs.

It is important to note that *ChemSource* is not meant to replace the chemistry curricula already in place. It is, as its name suggests, a resource to aid teachers in accomplishing their curricular objectives.

PHILOSOPHY OF *CHEMSOURCE*

The underlying philosophical principles that guided the development of *ChemSource* and that infuses each part with a unifying theme are:

- Teachers teach teachers. *ChemSource* is the pre-eminent example of teachers teaching teachers. Every part of *ChemSource* was planned, developed and written by active teachers.

- Cross-fertilization of ideas is essential to growth in chemistry teaching. Teachers cannot teach in isolation and remain effective. The best teachers are those who constantly learn from other teachers. *ChemSource* contains the best ideas of many active teachers from around the nation, providing a cross-fertilization possibility that teachers could previously obtain only by attending workshops and meetings.

- Hands-on activities are the most important activities. Chemistry teaching is one of the most important endeavors there is, and the most important aspect of chemistry teaching is hands-on activities, whether these be in the laboratory, in the classroom, or in the field. Every aspect of *ChemSource* has focused on these activities. Each *SourceBook* module sets forth content and its place in the curriculum; concept development through laboratories, demonstrations and discussion activities follow and take up the largest portion of each module; teaching tips to reinforce the activities then follow, but the centerpiece remains the hands-on activities in every case. Throughout *SourceBook,* laboratory activities are used inductively to introduce concepts, or deductively to verify concepts. The hands-on theme is echoed in the *SourceView* tapes. Every episode involves an activity of some sort, with an emphasis on hands-on laboratory and demonstration activities.

- Problem solving can/should be taught using higher order thinking skills. Another important philosophical underscore to the *ChemSource* project is the value of higher order thinking in the development of chemical concepts. *SourceBook* questions and follow-up activities, and *SourceView* problem-solving episodes emphasize the importance of encouraging higher order thinking for both students and teachers.

THE *SOURCEBOOK* MODULES

1. Basic Modules (written from a conceptual standpoint in which the chemical principles involved are developed and organized):
    - Acids and Bases (ACID)
    - Atomic Structure (ATOM)
    - Chemical Bonding (BOND)
    - Condensed States (COND)
    - Electrochemistry (ELEC)
    - Chemical Equilibrium (EQIL)
    - Gases (GASS)
    - Molecular Geometry (GEOM)
    - The Mole (MOLE)
    - Nuclear Chemistry (NUCL)
    - Organic Chemistry (ORGN)
    - Oxidation-Reduction (OXID)
    - Periodicity (PERD)
    - Solubility and Precipitation (PPTN)
    - Rates of Reaction (RATE)
    - Simple Chemical Reactions (RXNS)
    - Separations (SEPN)
    - Solutions (SOLN)
    - Stoichiometry (STOI)
    - Thermochemistry (THER)
2. Descriptive modules (similarities and differences in properties and behavior of elements and compounds are described and related to fundamental chemical principles):
    - Alkali Metals (ALKA)
    - Halogens (HALO)
    - Transition Elements (TRAN)
3. Enrichment/Applied Modules (presenting chemistry as a human activity, practiced by real people, with everyday life applications):
    - Biogeochemical Cycles: Carbon, Oxygen, Nitrogen (BICY)
    - Chemistry and Food (FOOD)
    - Enzymes: Biochemical Catalysts (ENZY)
    - Forensic Chemistry (FORS)
    - Industrial Inorganic Chemistry (INDL)
    - Instrumentation (INST)
    - Materials Science: Ceramics and Glasses (MATR)
    - Chemistry in Medicine (MEDI)
    - Photochemistry (PHOT)
    - Polymers (POLY)
    - Inorganic Qualitative Analysis (QUAL)
    - Chemistry of Rocks, Minerals and Gems (ROCK)
    - Chemistry of Seawater (SEAW)
4. General Materials
    - Preface (PREF)
    - Chemical Pedagogy (PEDA)
    - User's Guide (USER)
    - Introductory Module (INTR)

- Chemical Safety (SAFE)
- Teaching Chemistry to Persons with Disabilities (DISB)
- Chemistry in Your Career (CRER)
- Library Resources (LIBR)
- Professional Organizations (PROF)
- Cross-Referencing (CROS)
- Computers and Chemistry (COMP)
- Indices: Chemicals Index (INDC), Demonstrations Index (INDD), Experiment Index (INDE), Media Index (INDM), Name Index (INDN) and Subject Index (INDS)

5. *SourceBook* Editorial Policy

- *SourceBook* Periodic Table columns will not be numbered (see *Periodicity* module, Content-in-a-Nutshell)
- Electron configuration will be written in order of principal quantum level
- Transition elements are defined as the elements in the *d*- and *f*-blocks (see *Atomic Structure* module)
- Condensed formulas of organic compounds will be used throughout *SourceBook* unless ambiguity results (see *Organic Chemistry* module)

6. Definitions of Laboratory Activity Level

• Basic	Appropriate for use in classes oriented toward general chemistry literacy. Students probably have limited mathematical skills and are not preparing for scientific careers.
• General	Appropriate for use in classes for college-bound students who may or may not be science-oriented, but have adequate mathematical skills for data processing.
• Honors	Appropriate for use in classes of selected students whose mathematical background, science aptitude and interest are strong. These students are probably considering careers as science professionals.
• Advanced Placement	Appropriate for use in second-year chemistry courses which may or may not be based on AP curriculum. [Many of the honors laboratory activities could be used for AP review.]

A TRIP THROUGH A TYPICAL *SOURCEBOOK* MODULE

All *SourceBook* modules use the same organizational structure. The five major subheadings are:

- *Topic Overview*
- *Concept / Skills Development*
- *Links and Connections*
- *References*
- *Appendices*

(A sixth subheading: *Extensions and Projects* is included in many modules.)

The *Topic Overview* includes:

Content in a Nutshell - a capsulized coverage of the material in the module.

Place in the Curriculum - a statement describing how and when the topic can be integrated into the chemistry curriculum.

Central Concepts - a listing of central ideas, principles and definitions which form the basis of the module.

Related Concepts - concepts which support and relate to the central concepts.

Related Skills - skills needed by the student to carry out the module activities.

Performance Objectives - a description of what the student will be able to do after completing the activities and studying the other material in the module. Performance Objectives are a key component of effective lesson plans which tie content to student learning.

The *Concept / Skills Development* section forms the backbone of every module. Here, the user will find key activities (laboratory activities and demonstrations) which support the topic. *SourceBook* is both a laboratory book and a resource book contained in a single package.

Laboratory Activities include both the Student Version (which can be duplicated as a handout) and the Teacher Notes which provide all the information teachers need to conduct the laboratory and assess student learning. The laboratory activities in the *Concept / Skills Development* section include:

Student Version:

- Introduction
- Purpose
- Safety
- Procedure
- Data Analysis and Concept Development
- Implications and Applications

Teacher Notes:

- Major Chemical Concepts
- Level
- Expected Student Background
- Time
- Safety
- Materials
- Advance Preparation
- Pre-Laboratory Discussion
- Teacher-Student Results
- Answers to Questions
- Post-Laboratory Discussion
- Assessing Laboratory Learning

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Demonstrations
Demonstrations stimulate student interest in a topic and set the stage for new ideas or reinforce material previously covered. Multiple demonstrations are included in each module so the teacher can pick and choose which to include in a lesson. Demonstrations in the *Concept / Skills Development* section include:

- Purpose/Description
- Materials
- Safety
- Procedure

Group and Discussion Activities
Cooperative learning through groupwork is an effective technique for achieving conceptual learning and creative problem solving. All modules include *Group and Discussion Activities* in the *Concept / Skills Development* section. These include but are not limited to:

- Key Questions
- Counterintuitive Examples
- Analogies and Metaphors

In addition to activities, the *Concept / Skill Development* section includes the following subtopics:

Tips for the Teacher - tips from the language of chemistry as well as common misconceptions.

History - historical development of the topic or biographical sketches of key individuals who made significant contributions.

Humor - cartoons, puzzles, drawings, anecdotes, puns, word search and crossword puzzles.

Media - commercially available films, videotapes, computer software, *etc*.

Links and Connections are a third subtopic found in each module. Connections are:

Within Chemistry

Between Chemistry and Other Disciplines

To the Contemporary World - these include links to personal experiences which relate to the topic as well as links to the community, suggesting field trips and knowledgeable individuals who could serve as a resource.

References make up the fourth subheading:
References cited in every module are written in a format which will facilitate interlibrary loans.

The *Appendix*, found at the end of each module, includes ancillary material, activities, transparency masters, *etc*. The latter are ready for duplication.

CROSS-LINKS BETWEEN MODULES

Some of the applied/enrichment modules include extensive background notes which are not likely to be found easily in standard chemistry sources. Many modules include information on specific equipment and instrumentation useful in studying the topic. Finally, a majority of the modules give ideas for extensions and research projects which could be adapted for Talented and Gifted or independent study projects.

While each module in *SourceBook* covers a particular topic, there are many links and connections between topics in other modules which can be located using the cross-referencing index. For example, while acids and bases are the subject of the module by that name, there are extensive links to acids and bases in many other modules. These links are outlined in the *Cross-Referencing* module (CROS).

Some educators consider Concept Maps a helpful pedagogical tool for review and reinforcement. While each individual may construct a different concept map for each topic, we have provided examples both below (on *Chemical Bonding*) and in the *Chemical Pedagogy* module (on *Acids and Bases*). Use of concept maps for review and assessment is also illustrated in the *SourceView* videotapes. Concept maps are most helpful when a concept is illustrated in a box or oval, and the link to other concepts in the overall topic are shown not only by connecting lines but by explanatory phrases.

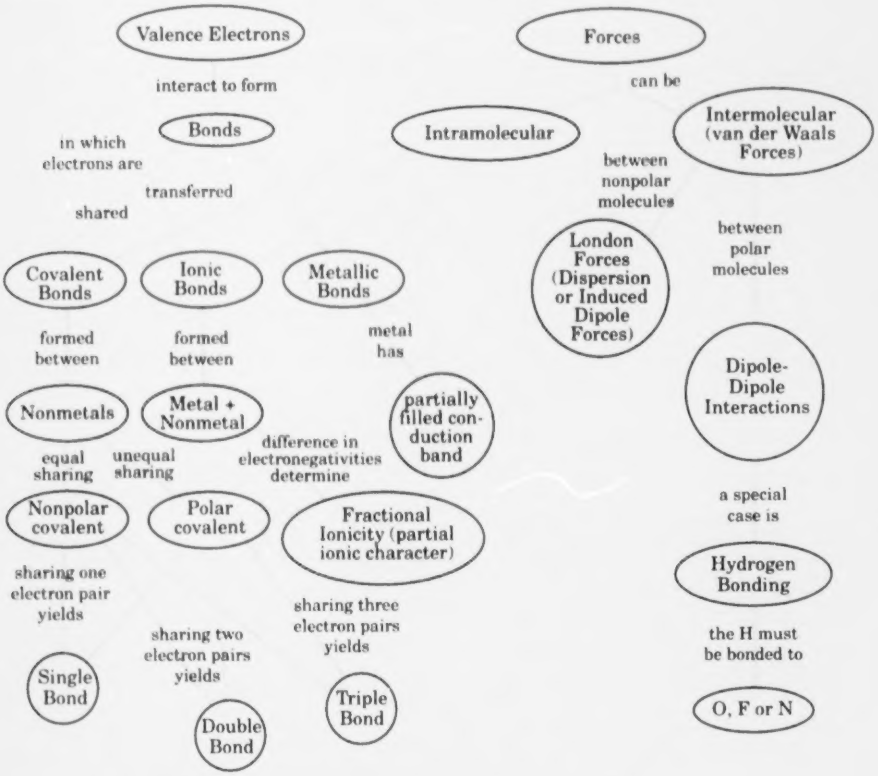
Valence Electrons
interact to form
Bonds
in which electrons are
transferred
shared
Covalent Bonds
Ionic Bonds
Metallic Bonds
formed between
formed between
metal has
Nonmetals
Metal + Nonmetal
partially filled conduction band
equal sharing
unequal sharing
difference in electronegativities determine
Nonpolar covalent
Polar covalent
Fractional Ionicity (partial ionic character)
sharing one electron pair yields
sharing two electron pairs yields
sharing three electron pairs yields
Single Bond
Double Bond
Triple Bond
Forces
can be
Intramolecular
Intermolecular (van der Waals Forces)
between nonpolar molecules
between polar molecules
London Forces (Dispersion or Induced Dipole Forces)
Dipole-Dipole Interactions
a special case is
Hydrogen Bonding
the H must be bonded to
O, F or N

Figure 1. Concept Map for Chemical Bonding

BUILDING A LESSON PLAN USING *SOURCEBOOK*

There are several approaches to developing lesson plans which use *SourceBook* as either a primary or secondary resource. A typical topic, acids and bases, illustrates how *SourceBook* materials can be integrated into the curriculum. First, a unit plan is presented which incorporates both the regular chemistry curriculum and *SourceBook* supplementary material into a two-week unit. Two detailed daily lesson plans follow, each using *SourceBook* material in a different way. The first plan uses the module *Acids and Bases* as the primary resource for the lesson. The second plan uses the standard chemistry curriculum for teaching central concepts, and supplements with enrichment material from several *SourceBook* modules.

The daily lesson plans presented here combine course content with up-to-date teaching strategies which make use of learning theory and research findings to promote student learning. Many schools are currently using similar lesson plan formats, such as Instructional Theory Into Practice (ITIP), which include a series of teaching steps. The steps are viewed as essential to thorough lesson development. The model for the daily lesson plan used on the following pages was selected because the steps incorporated have become guidelines for evaluating teacher performance in many teacher appraisal programs currently in use. The following elements are found in the daily lesson plans:

Focus or Anticipatory Set - means of generating student interest in the lesson that follows which involves all students and builds on previous knowledge.

Performance Objectives - description of what the student will be able to do after instruction.

Explanation - techniques used by the teacher to present the information.

Modeling - ways to demonstrate the learning process visually.

Checking for Knowledge and Understanding - assessing students' understanding of the learning using key questions, *etc.*

Monitoring and Adjusting - monitoring the progress of the learning and adjusting instruction to correct misconceptions.

Guided Practice - guiding students through the problems, skills or work to be done and providing feedback to the student.

Independent Practice - student practice without teacher assistance.

Enrichment - higher level activities provided to students which extend the knowledge of the subject.

Closure - activity which brings the lesson to a close and requires the student to summarize and internalize what has been taught.

SAMPLE UNIT LESSON PLAN

Acids and Bases

Day 1 Classifying Substances as Acids, Bases, Salts [See *SourceBook,* Acid/Base module: Activity 1]

Day 2 Classifying Substances [Activity, continued]

Day 3 Conceptual Definitions and Strength of Acids and Bases [See Daily Lesson Plan using *SourceBook* as the **primary** resource]

Day 4 pH and Indicators [See Daily Lesson Plan using *SourceBook* as a **secondary** resource]

Day 5 Soil pH [See *SourceBook, Industrial Inorganic Chemistry* module: Activity 1, Part 3 and Demonstration 5]

Day 6 Neutralization, Buffers, and Hydrolysis [See *SourceBook, Acids and Bases* module: Extensions]

Day 7 Neutralization [See *SourceBook, Acids and Bases* module: Suggestions for Other Laboratory Activities - Neutralization]

Day 8 Titration [See *SourceBook, Acids and Bases* module: Suggestions for Other Laboratory Activities - Titration of an antacid]

Day 9 Review of Concepts involving Acids and Bases

Day 10 Assessment or Test on Acids and Bases

Since acids and bases are typically covered in the second semester, the time spent on this topic can be shortened or lengthened depending on the material yet to be covered in the first year. Many excellent films, videos, and environmental considerations are often used to extend the coverage on the topic of acids and bases. An important link between acids, bases and periodicity is found in the *Periodicity* module, Activity 2: Trends in Oxy-hydrogen Compounds of Period 3 Elements. If this activity was not used during the study of periodicity, it could be used as a logical laboratory extension when studying acids and bases.

DAILY LESSON PLAN USING *SOURCEBOOK* MODULE: *ACIDS AND BASES* AS THE PRIMARY RESOURCE

Acids and Bases
Day 3: Conceptual Definitions and Strength

Previous assignment due today: *SourceBook* Activity 1: Classifying Substances Based on their Reactions

Homework assignment - Due date:

Materials and equipment needed:

- 2 Burets
- 2 Ringstands
- 2 Clamps
- 2 Beakers, 250-mL
- Conductivity apparatus (with 60-W light bulb if possible or small-scale apparatus with piezoelectric buzzer)

0.01 M or 0.1 M Sulfuric acid, H_2SO_4
0.01 M or 0.1 M Barium hydroxide, $Ba(OH)_2$
Phenolphthalein solution, 1% in 95% ethanol
Stirring rod or magnetic stirrer
Transparency Masters

OUTLINE OF SUBJECT MATTER AND STRATEGIES FOR REACHING TODAY'S OBJECTIVES

1. **Focus or anticipatory set**
   a. Perform Demonstration 4, Part 2. Tie in demonstration with Activity 1 performed yesterday.
2. **Performance objectives**

(*SourceBook* Performance Objectives 1, 7, 8 and 11)

The student should be able to:

   a. define and classify acids and bases operationally and conceptually
   b. distinguish between a strong acid (or base) and a weak acid (or base)
   c. present useful mental pictures of ionization, dilute/concentrated and weak/strong acids and bases
   d. express opinions about the role of acids and bases in daily life.

3. **Content**
   a. Central Concepts
      1) Conceptual definitions (#2). Include: History of Arrhenius, Brønsted and Lowry.
      2) Weak and strong acids/bases (#5, 6)
   b. Related Concepts
      1) Ions (#1)
      2) Classifying acids and bases (#3)
   c. Links and Connections
      1) Within Chemistry
      2) Between Chemistry and Other Disciplines
4. **Modeling**
   a. Transparency Masters of Pictures in the Mind
      1) Ionization
      2) Concentrated/Weak; Dilute/Strong Acids and Bases
   b. Metaphors and Analogies
5. **Checking for Knowledge and Comprehension**
   a. Key questions 1, 2, 5-7.
6. **Monitoring and Adjusting**
   a. Common Student Misconceptions. See 1-4 and 8.

7. **Guided Practice**
    a. Students asked to describe acids and bases in operational and conceptual terms.
    b. Student asked to compare and contrast the following terms with reference to acids and bases: concentrated/weak; strong/dilute.
    c. (Advanced) Students asked to relate strength of acids/bases in terms of ionization constant expressions, K_a and K_b.
8. **Independent Practice**
    a. Student will answer teacher-assigned questions on acids and bases which are found in any chemistry textbook, or
    b. Student will prepare a list of 15-20 products from a grocery store which includes an acid or base as one of the ingredients listed on the label, or
    c. Student will write a one-page paper on a specific link of acids and bases to another discipline. Example: environmental problems with acid rain.
9. **Enrichment**
    a. Student might conduct library research on G.N. Lewis and/or refer to most chemistry texts for the Lewis definition of acids and bases. This more encompassing definition could be compared with the Arrhenius and Brønsted-Lowry definitions.
10. **Closure**
    a. Students summarize definitions of acids and bases. [Operational definitions should be reviewed from the previous day's laboratory activity; conceptual definitions should be summarized from today's topics.]
    b. Transparency Masters of Humor dealing with acids and bases.

[Note: The activities listed above can be performed in a typical class period. Activities 7(a), 7(b) and 10 could be used to review the material on the following day. Activities 8(c) and 9 are more long-term and could constitute a unit assignment.]

DAILY LESSON PLAN USING *SOURCEBOOK* AS A SECONDARY RESOURCE

Acids and Bases
Day 4: pH and Indicators

Previous assignment due today: Teacher-assigned questions on acids and bases generated from a chemistry textbook.

Homework assignment - Due date:

Materials and Equipment needed:
Natural indicators: Red cabbage juice, grape juice, *etc.*
Vinegar
Baking soda (sodium bicarbonate), $NaHCO_3$
6 Petri dishes, slant test tube apparatus, or 50-mL beakers
Transparency masters

OUTLINE OF SUBJECT MATTER AND STRATEGIES FOR REACHING TODAY'S OBJECTIVES

1. **Focus or anticipatory set**

   Consider the flowering shrub, hydrangea and how the flowers can be red or blue depending on the nature of the soil. [*SourceBook: Acids and Bases* module, Content in a Nutshell]

2. **Performance Objectives**

   The student should be able to:

   a. classify a solution as acid, base or neutral, given the hydrogen-ion or hydroxide-ion concentration.

   b. calculate the hydrogen-ion or hydroxide-ion concentration in an aqueous solution.

   c. define pH and perform calculations involving pH.

   d. explain why indicators change color.

   e. select appropriate indicators for acid/base solutions.

3. **Content**

   a. Central Concepts [Source: Most chemistry textbooks]
      1) Hydrogen-ion and hydroxide ion-concentration of acids and bases
      2) pH
      3) Acid-base indicators

   b. Related Concepts [Source: *ChemSource*]
      1) pH of seawater [*SourceBook: Chemistry of Seawater* module, Activity 1, part D]
      2) Qualitative Analysis and pH [*SourceBook: Inorganic Qualitative Analysis*, Links and Connections within Chemistry]
      3) Household substances and pH [*SourceBook: Food and Chemistry*, Tips for the Teacher: Shampoo and Soda Beverages]

4. **Modeling** [Taken from *SourceBook: Acids and Bases* module]

   a. Demonstration 2: Nature of Indicators

   b. Metaphors and Analogies, #1

   c. Transparency Masters
      1) Language of Chemistry: Indicators and pH; pH of Common Substances

      2) Pictures in the Mind #3 and #4

5. **Checking for Knowledge and Comprehension**

   a. Key Questions for Discussion
      1) How is an acid or base identified on the basis of the hydrogen-ion or hydroxide-ion concentration?
      2) What is pH and how is it determined?
      3) How can pH be used to classify common substances?
      4) What are indicators and how are they used in acid/base chemistry?

6. **Monitoring and Adjusting**
   a. Student will interpret the results of tests on unknown substances using indicators.
   b. Common Misconceptions #5, 6, 7 [From *SourceBook* module: *Acids and Bases*]
7. **Guided Practice** [See mathematical calculations involving pH in *SourceBook:* General Resource module]
   a. Students asked to classify a list of substances as acids or bases when hydrogen-ion and hydroxide-ion concentration is known.
   b. Students asked to calculate the pH of solutions if the hydrogen-ion and hydroxide-ion concentration is known.
8. **Independent Practice**
   a. Student will answer assigned questions on pH and indicators which are found in most chemistry textbooks, or
   b. Given strips of pHydrion™ paper, the student will determine the pH of 10 common substances located in the home or workplace, or
   c. Student will prepare a natural indicator from flower or fruit concentrate at home and test it on common acids and bases.
9. **Enrichment**
   a. Student will investigate the top chemicals used in the U.S. and determine how many are acids and bases. [*SourceBook: Industrial Inorganic Chemistry* module, Appendix]
   b. Student will write a research paper on commercial processes for the production of common acids and bases like the Ostwald, Solvay, and Haber-Bosch processes. [*SourceBook: Industrial Inorganic Chemistry* module]
   c. Student will conduct a research project on pH of lakes and streams in a given area and relate results to the presence or absence of pollution.
10. **Closure**
    a. Students summarize the relationship between acids and bases, pH, and indicators
    b. A pH rhyme [*SourceBook: Acids and Bases* module, Tips for the Teacher #5]
    c. Transparency Masters of Humor dealing with pH [*SourceBook: Acids and Bases* module, Humor]

The authors of *ChemSource* have provided a comprehensive resource for chemistry teachers. As every science teacher knows, however, new information, techniques, and strategies will continually impact the way chemistry is taught. While the materials found in the *SourceBook* modules were developed by thirty-six teachers across the U.S., a tremendous amount of experiential knowledge still resides in the thousands of chemistry teachers around the country who have developed their own bags of tricks, resources, *etc.* We welcome the contributions, suggestions, questions, and ideas from all chemistry teachers who use *ChemSource* to extend the process of teachers helping teachers. A feedback form has been provided below. Please take time to respond. Chem*Source* will never be a finished product. Periodic revisions and supplements will be necessary to continue the resource development for new and old chemistry teachers alike. Your contribution to this process is both encouraged and welcomed.

The User's Guide was developed by Beatrice Epperson and Mary Virginia Orna.

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(Please photocopy a separate form for each *SourceBook* section)

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We invite you to become part of an on-going review team by completing a facsimile of the form below and returning to us your corrections and assessments to be incorporated in the next edition of *SourceBook*. We also welcome additional humor, anecdotes, games, links and connections, as well as any teaching aids. For your effort, you will be credited for your contributions in the next edition.

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COMMENTS: Did you find this *SourceBook* module or chapter useful? Why or why not? Was the material: too complicated? too simple? Were the explanations sufficient? Please comment. Did the laboratory activities and demonstrations "work?" If not, what was the problem?

IMPROVEMENTS: What additions, deletions or alternatives would improve this *SourceBook* module/chapter? Please comment.

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ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

INTRODUCTORY: GENERAL RESOURCE

Practical Aspects of Teaching Chemistry including Measurement Skills, Unit Conversions, Uncertainty in Measurement, Problem Solving,and Chemical Mathematics

Practical Aspects of Teaching Chemistry including Measurement Skills, Unit Conversions, Uncertainty in Measurement, Problem Solving, and Chemical Mathematics

Measurement

Use of Units

The *metric system*, based on meter and kilogram units, was first adopted in France in 1790. In common usage throughout Europe, U.S. residents have resisted its use in daily commerce although it was legalized by the Congress in 1866, and declared the U.S. fundamental standard in 1893. In 1960, the eleventh International Conference on Weights and Measures proposed major changes in the metric system and suggested the name *International System of Units* (abbreviated SI) for this modernized metric system. Use of SI is encouraged by the U.S. Metric Conversion Act of 1975 but, although the world's major countries have agreed to adopt SI for scientific as well as everyday use, no U.S. law has ever mandated its exclusive use.

Base Units

SI consists of seven base units (Figure 1). All other SI units are derived from these seven.

Quantity Measured	Unit Name	SI Symbol
Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Electric current	Ampere	A
Temperature	Kelvin	K
Amount of substance	Mole	mol
Luminous intensity	Candela	cd

Figure 1. SI base units.

A prefix system representing various powers of 10 has been defined in SI. Each base unit can be enlarged or reduced to a convenient size. Figure 2 lists some commonly used prefixes.

Prefix	Symbol	Basic Unit Amount	Exponential Notation
mega	M	1 000 000	10^6
kilo	k	1 000	10^3
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	µ	0.000 001	10^{-6}
nano	n	0.000 000 001	10^{-9}
pico	p	0.000 000 000 001	10^{-12}

Figure 2. Prefixes for fractions and multiples of SI units.

For example, the distance between two towns is better expressed as 9.5 km rather than 9500 m. The distance between hydrogen and oxygen nuclei in a water molecule is about 0.000 000 000 097 m. Obviously this distance is better expressed as 0.097 nm or 97 pm. Indeed, atomic dimensions are of the order of a fraction of a nanometer or 10s of picometers. Thus, the unit chosen is the one most appropriate for the magnitude of the measurement.

The above examples also illustrate the usefulness of scientific, or exponential notation, used to express very small and very large numbers. In this notation, any number may be expressed as a number between 1 and 10 times the proper exponent of 10. The hydrogen-oxygen distance in the previous paragraph thus becomes 9.7×10^{-11} m.

To cite another example, 1 g of water is known to contain about 33 460 000 000 000 000 000 000 water molecules. In scientific notation, this becomes 3.346×10^{22} molecules. In scientific notation, a number is expressed in the form $n \times 10^p$ where $10 > n \geq 1$ and p is any integer (whole number). Often it is convenient to write a number in a format such that all values of p are the same to simplify comparison. This is not scientific notation, but simply a form of exponential notation, which by definition does not restrict the range of n.

Derived Units

All other SI units can be derived from the seven base units described above. A familiar example from physics is the unit of force (the newton N), derived from Newton's second law of motion, force = mass x acceleration, or F = ma. Since acceleration is defined as change of velocity per time, its units are meters/second/second or meters/second^2. This can be written as m/s^2 or m s^{-2}. Thus, force is kg m s^{-2}, which is defined as the newton.

Pressure is defined as force per unit area, or newtons per square meter. Its units are therefore N/m^2, or N m^{-2}. One newton per square meter is defined as a pascal (symbol Pa), the derived SI unit of pressure. Since the pascal is a very small unit, roughly equal to the pressure exerted by a dollar bill lying flat on a table, a more convenient unit, the bar, is defined as 10^5 pascals. This is the approximate pressure exerted by our atmosphere at sea level and is expressed as one atmosphere of pressure. Although in common usage in chemistry, the term atmosphere is no longer acceptable with SI units.

The SI derived unit of volume, m^3, is still not widely used because the liter (L) and milliliter (mL) are so common. All laboratory glassware is still calibrated in L or mL, and these units will continue in common usage. Figure 3 lists certain derived SI units, other widely used units, and their SI equivalents. You will also encounter "obsolete" units and symbols in texts. Examples are the Angstrom (Å = 10^{-8} cm), micron (μ = 1 micrometer), and extinction coefficient [ε = absorptivity (a)].

Quantity	Unit name	Symbol	SI Derivation	Non-SI unit	SI equivalent
Acceleration			m s^{-2}		
Area			m^2		
Celsius temperature				Degree Celsius (°C)	0 °C = 273.15 K
Density			kg m^{-3}	g L^{-1}	1 g L^{-1} = 1 kg m^{-3}
Electric charge	Coulomb	C	A s		
Electric potential	Volt	V	J C^{-1}		
Energy	Joule	J	N m	Calorie (cal)	1 cal = 4.184 J
Force	Newton	N	kg m s^{-2}	Dyne	1 dyne = 10^{-5} N
Frequency	Hertz	Hz	s^{-1}		
Molar volume		V_m	m^3 mol^{-1}	L mol^{-1}	1 L mol^{-1} = 10^{-3} m^3 mol^{-1}
Power	Watt	W	J s^{-1}		
Pressure	Pascal	Pa	N m^{-2}	Atmosphere	1 atm = 760 torr = 760 mm Hg = 101325 Pa
Volume			m^3	Liter (L)	1 L = 10^{-3} m^3 = 1 dm^3

Figure 3. Derived SI and commonly used equivalent units.

Use those units most familiar to you and easiest for your students to visualize. Remember, all units are equally unfamiliar to most students. At the very least, use SI where possible, justifying substitute units when more convenient.

MEASUREMENT IN THE LABORATORY

Most laboratory manuals have an introduction or appendix that details proper use of balances, graduated cylinders, pipets, burets, *etc.* Be sure students know how to use each instrument they will manipulate in the laboratory.

Laboratory instruments such as balances and pH meters come with detailed instructions. A file should be created for each such instrument to assure availability of operating instructions, maintenance procedures, service history, and warranty information. Samples of previous results from the instrument are also useful teaching aids and should be kept in this file. Issuing photocopies of the instrument manual rather than originals to students is advised.

Balance

Weighing substances is potentially the most precise measurement used by chemists. Measurement is done by **counter balancing** the sample of unknown mass against objects of known mass. This balancing act is accomplished by means of a delicate pivot point called a knife edge. Thus the measuring instruments used are called **balances** (Figure 4).

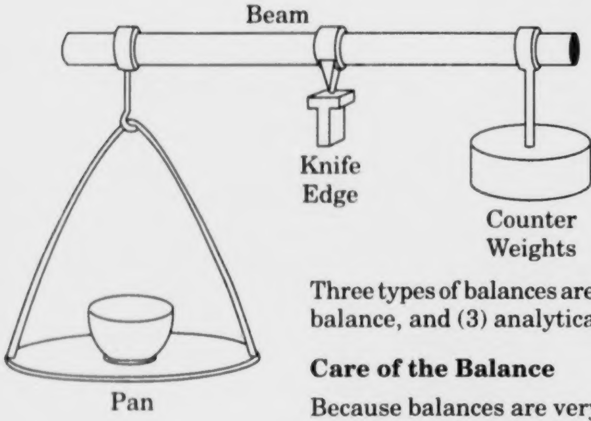


Figure 4. Schematic of a balance.

Three types of balances are discussed: (1) triple beam balance, (2) top-loading electronic balance, and (3) analytical electronic balance.

Care of the Balance

Because balances are very delicate and expensive instruments, the following rules of care must be followed:

1. Never place substances directly on the balance pans as this may lead to corrosion and eventual malfunctioning. Beakers, paper "boats," *etc.* should be used to hold the substance being weighed.
2. With a spatula or by pouring (see Figure 8) remove a substance from a container and place it on a balance pan.
3. If a spill occurs, use a balance brush to gently sweep it away. (Be sure the electronic balance is shut off before you start cleaning.) If the spill involves more than a few grains of material or involves a highly corrosive substance, provide clean-up instructions for students.
4. Do not move a balance without locking the working mechanism; otherwise, you risk upsetting the leveling and the machinery may jam. You also risk damaging the knife edge.

Use of the Analytical Electronic Balance

The analytical balance (Figure 5) discussed here is the older, partially mechanical version still being used in many laboratories. This balance appears electronic and foolproof because all that shows is a balance pan, some knobs, and a mass read-out window. However, it is basically like other balances; the knife-edge, balance beam, and counterweights are enclosed in the housing. The balance is very precise (± 0.001 to ± 0.0002 g) making it also very delicate and very easily damaged. It must be used with care.

In general, the electronic balance has four types of controls: zeroing knob, large weight knobs (turning these causes weights to be placed on or off the balance beam), fine weight knob (for a better estimate of milligram weights), and on-off lever with three positions: Off or Full Arrest (balance beam held in place), Partial Arrest (balance beam has limited movement), Full On (balance beam has complete freedom of movement.

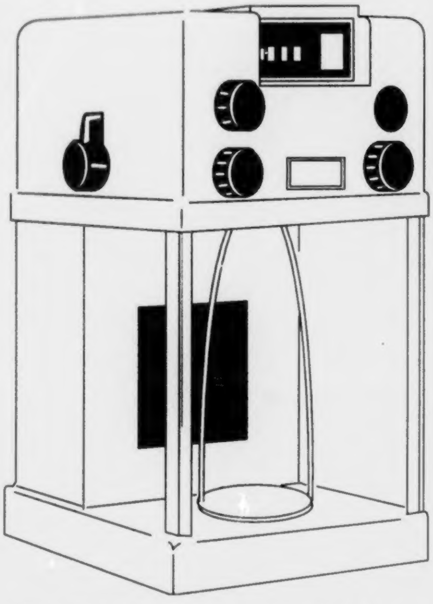


Figure 5. Analytical electronic balance.

Zeroing

Follow these steps:

1. Check that the instrument is **Off**. Adjust all weight knobs to zero; close doors.
2. Turn **Full On** and adjust to read 0.0000 g by gently turning the Zeroing knob.
3. Turn **Off**.

Weighing

Follow these steps:

1. Place the container to be weighed on the pan. The instrument must be **Off** whenever placing things on or removing them from the pan.
2. Turn lever to **Partial Arrest** position.
3. Gently and slowly adjust the 10 g knob until the projected scale moves abruptly upwards. Then turn the knob back one step.
4. Repeat Step 3 using the 1 g knob and then the 0.1 g knob if there is one.
5. Turn the lever gently to the **Off** position and then to the **Full On** position. Be sure the doors are closed.
6. When the projected scale stops moving, adjust the **Fine Weight** knob.
7. Read and record the mass. (Some balances, particularly those of European make, use a comma as a decimal point.) Be sure final mass is taken with the lever in the **Full On**, not the **Partial Arrest** position.
8. Turn **Off**.

Use of Top-Loading Electronic Balance

The newest, totally electronic, version of this balance (see Figure 6) uses a pressure sensitive strain gauge to determine mass. The object's mass affects the electronic signal passing through the gauge; this signal is displayed as a digital readout. The instrument can be as precise as many analytical balances (± 0.001 g) and is easier to use. Top loading balances are becoming the instrument of choice for routine laboratory work. However, they are delicate and expensive to repair. Thus they must also be used with care.

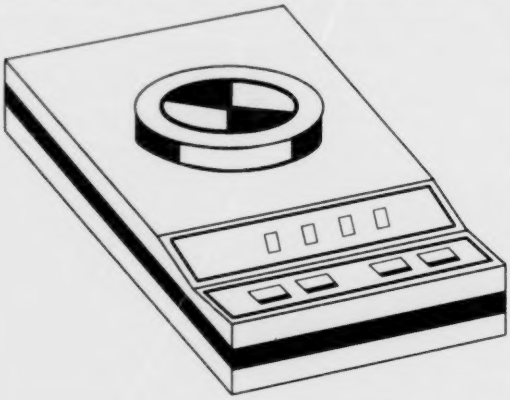


Figure 6. The top-loading balance.

The controls, in general, are as follows: **On-Off** switch, **Calibration** switch, **Tare** switch. Some balances are different, so be sure to read the instruction manual accompanying the balance.

Weighing

Follow these steps:

1. Turn on the instrument using the **On-Off** switch.
2. Push the **Calibration** switch. The digital display will rapidly run through a series of displays as the electronics are calibrated. When the digital display reads 0.000 g, the calibration is complete.
3. Place the object to be weighed on the pan, wait a few seconds for the digital display to stabilize, then read mass of the object in grams.
4. Turn the instrument **Off** after weighing.

Taring

The **Tare** switch allows the display to be set to 0.000 g even with an object on the balance pan. It is used when the mass of the substance and not the mass of the substance and container is needed. Use the **Tare** feature only when the mass of the container is not needed.

1. Weigh the container as described above. After the display settles down, push the **Tare** switch. The display should read 0.000 g.
2. Add the desired substance to the container. The display now shows the mass of the added substance.

Use of the Triple Beam Balance

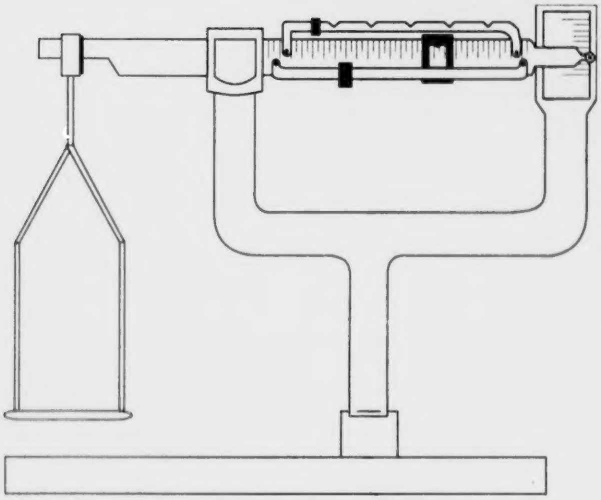


Figure 7. The triple beam balance.

All the essential parts of this balance are visible (Figure 7): knife edge, balance beam, and counterweights (technically called *poise*). This balance can be precise to ± 0.01 g if handled properly. Readings may be interpolated to ± 0.001 g, which probably cannot be repeated, but is a good exercise for students. Be careful not to move the balance arm sideways (by bumping it) as this might move the knife edge out of position.

Zeroing

Follow these steps:

1. With the pan empty, set all counterweights to their zero position. Be sure your fingers are clean and dry before handling the weights. Use a spatula or a pencil's eraser to move the 0.1 g sliding counterweight.
2. When zeroed, the pointer will swing an equal distance either side of the center line or will rest at the center line. If not, adjust the screw nuts at the extreme left of the balance.

Weighing

Place the container to be weighed on the pan. Using the instructions given in "Zeroing," move the counterweights until the pointer again rests on or makes equal swings about the center line. Be sure the larger weights are properly seated in the notches on the arms. Read and record the total mass and then return the weights to their zero positions.

Tips on Precise Weighing

To obtain the most precise (reproducible) results from weighing, the following steps should be observed. These are particularly important when using the high precision analytical balance.

1. Check the leveling bulb on the instrument. If it is not dead center, make the needed adjustments, usually adjusting the heights of one of the legs on the balance's underside.
2. Always use the same balance in weighing for a particular experiment. This will allow small constant errors in the balance to cancel in the final calculations.
3. Recheck the zero adjustment of the balance occasionally.
4. Be sure the balance and the area around it are kept spotlessly clean.
5. Never use fingers to directly handle a container being weighed. Fingerprints weigh a few milligrams each and evaporate slowly. Handle containers with tongs or with a lint free paper band as shown in Figure 8.

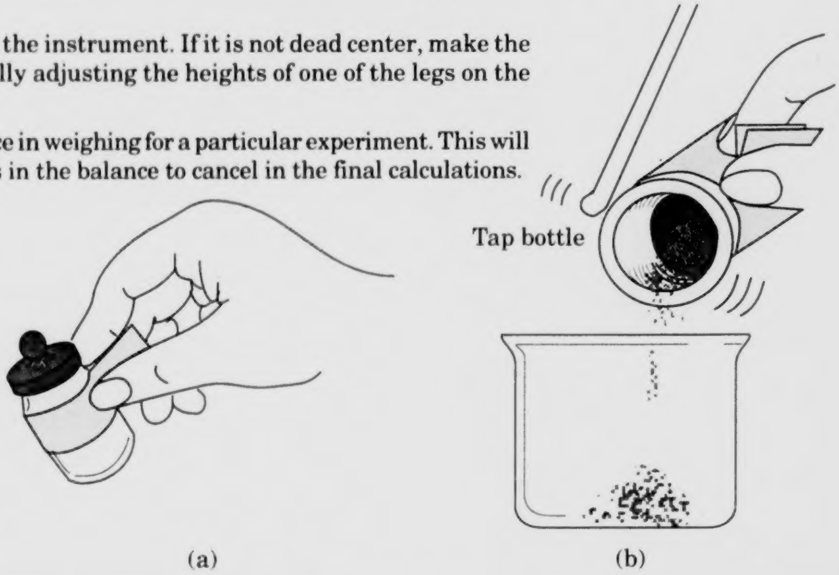


Figure 8. Transferring substances. (a) Handling a weighing bottle with a paper band. (b) Transfer procedure.

Tips on Balance Care

Balances do not stay zeroed forever. Triple beam balances usually have metal shot in the pan holder, which corrodes with time. The shot changes weight and needs to be removed, washed, dried, and supplemented every few years; otherwise, the balance will lose zeroing ability via the screw adjustment. Number balances and their removable pans, preferably by engraving, to improve the chance of keeping them together. Mixing pans and balances generally produces an unzeroable balance. Electronic balances require frequent recalibration according to their instructions. Servicing by a factory representative every few years is a good investment. You can save money by assembling balances from several departments or schools at a central place for a single service call.

Volume Measurements

Several types of glassware may be used to measure liquid volumes. Some give only approximate measurements (droppers, graduated beakers, and flasks), while others give very precise measurements (graduated cylinders, pipets, burets, volumetric flasks). When using more precise glassware one needs to know how to read a meniscus and if the glassware is designed *to contain* or *to deliver* the measured volume.

The liquid surface in any tube will be slightly curved, not perfectly horizontal. This curved liquid surface is called a *meniscus*. It will be curved downwards (as with water) when the attraction of the liquid molecules to the atoms of the glass is stronger than the attraction of the liquid molecules for each other. With liquids such as mercury, the meniscus is curved upwards because the mercury atoms experience a stronger attraction for each other.

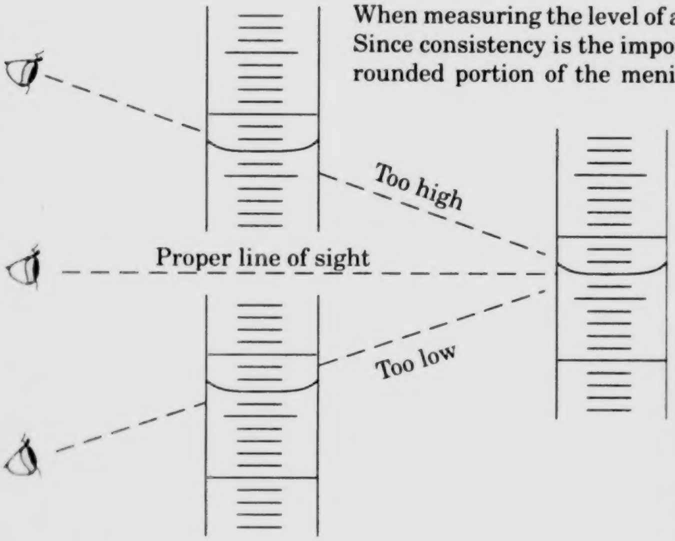


Figure 9. Reading a meniscus level.

When measuring the level of a liquid, decide which part of the meniscus to read. Since consistency is the important consideration, chemists read tangent to the rounded portion of the meniscus (*i.e.*, the bottom of the meniscus for water solutions). When reading a liquid level, the eye must be *on the same horizontal level* as the meniscus to avoid the problem of parallax, that is, choosing the wrong measuring line due to an improper viewing angle (see Figure 9).

Precise volumetric glassware is marked to measure either the volume of liquid contained or the volume of liquid that can be delivered (extracted). Such containers are labelled TC (for *to contain*) and TD (for *to deliver*), respectively. No single volume marking can precisely measure both of these quantities because all the liquid cannot be poured from a container. Some will always remain behind, clinging to the container walls. Be aware of which glassware is marked to measure liquid held or liquid delivered and use accordingly.

Droppers and Graduated Beakers

These types of glassware are used to give approximate volume measures. The dropper (often called medicine dropper or dropping pipet) is the quickest way to measure volumes of less than a few milliliters. The measurement is made by counting free, *vertically* falling drops assuming 1.0 mL = 18 drops. The method is precise to ± 10 % as one mL may be equal to 16 to 20 drops depending on the particular dropper used. Since dropper bores are not identical, it is necessary to calibrate the dropper before use. Count the number of drops required to fill a 10-mL graduated cylinder to the 1 mL line or weigh some number of drops of water and convert grams to milliliters (1 g = 1 mL). The angle of the dropper to the horizontal affects drop size. Calibrate and use at the same angle. For larger volumes, the markings on beakers and flasks can be used. These are accurate to only ± 5 % because of the large diameters of the glassware and the manufacturing methods. It is not important here whether the volume marking refers to the volume held or delivered since the inherent error is so large.

Graduated Cylinders

The volume markings measure the volume of liquid contained (or delivered) by the cylinders. Their accuracy is about ± 1-2 %. The error is much larger if the cylinder is used to measure liquid delivered. One way around this problem is to use a cylinder marked TD. The graduated cylinder should be calibrated ahead of time. For example, a graduated cylinder is to deliver 100 mL of solution several times in an experiment. Take two 100-mL cylinders and fill one with water to the 100.0 mL mark. Now pour this water carefully into the second, dry cylinder. Read the volume delivered. Within the error limits of ± 1-2%, this is the volume the first cylinder will deliver throughout the experiment assuming the pouring procedure to be the same during the experiment.

Pipets and Bulbs

Pipets deliver measured volumes with an uncertainty of only about ± 0. 01 % and should be used for highly accurate work. There are several types of pipets as shown in Figure 10. Measuring pipets deliver any volume up to a maximum capacity; volumetric or transfer pipets deliver a fixed volume. The pipetting bulb is used to draw the liquid into the pipet. (There is the temptation to disregard the bulb and draw the liquid into the pipet by mouth suction. *Do not give in to temptation!* Always use the bulb! Mouth suction can easily lead to ingestion of corrosive or poisonous substances.) The proper use of the pipet and bulb requires a fair amount of skill. It is worth practicing to perfect your skills using distilled water. You may find that the use of expensive bulbs with valves is worth the extra cost in time saved and mistakes avoided.

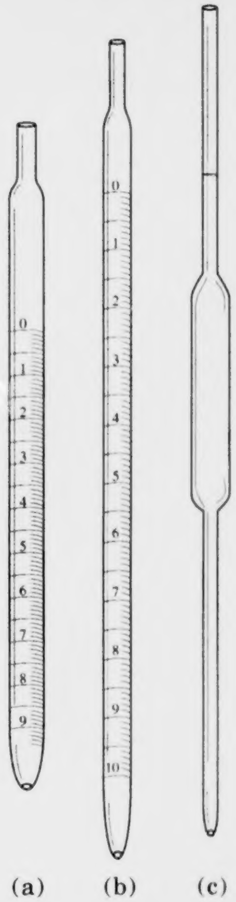


Figure 10. Pipets. (a & b) Measuring pipets and (c) volumetric or transfer pipet.

Preparing the Pipet

In order not to contaminate or dilute the solution being measured, the pipet must be cleaned and the inside surface coated with the solution to be measured.

1. Draw distilled water into the pipet with a bulb until about 1/5 of its "fat" portion is filled. Take off the bulb while quickly placing your forefinger over the pipet end. This will keep the water from draining out of the pipet. Keeping your finger over the pipet end, tip the pipet to a horizontal position. Rotate the pipet so that the entire "fat" portion is re-peatedly coated with water (Figure 11).

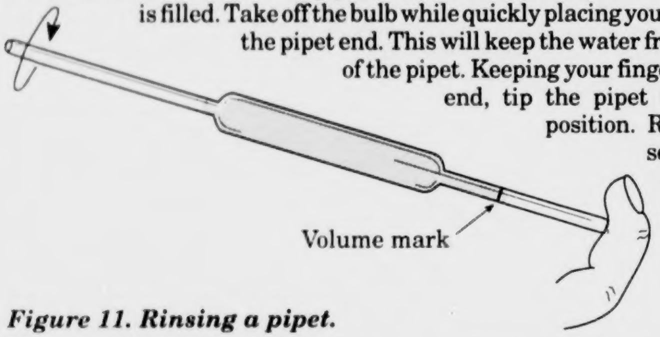


Figure 11. Rinsing a pipet.

Now allow the water to move down the "bulb" stem until it is about 2/3 of the way between the volume mark and the end of the stem. This is done by tipping the pipet to about a 45 degree angle and slightly releasing your finger pressure. (It is wise to not let any liquid reach the end of the "bulb" stem. If the end is wet it will be almost impossible to control the flow of liquid with finger pressure during the actual measurement.)

Tip the pipet back to the normal vertical position ("bulb" stem up) and let the water drain into a sink. Check the coating of the inside surface of the pipet. If it is uniform, with no breaks or localized droplets formed, proceed to Step 3.

2. An uneven coating of water on the inner surface of the pipet indicates that some insoluble materials, such as "oils," are adhering to the glass. Wash the pipet two to three times with a mild soap solution using the procedure outlined in Step 1. Then rinse out the soap solution with four to six treatments of tap water.

3. Using the procedure outlined in Step 1, rinse the pipet twice more with distilled water. Then rinse it twice with the solution to be measured. This will replace the water coating with a coating of the solution.

Measuring Procedure

The inside of the pipet should be coated with the solution to be measured as described above. If the same pipet is used to measure several solutions of the same substance at different concentrations, start with the one of lowest concentration.

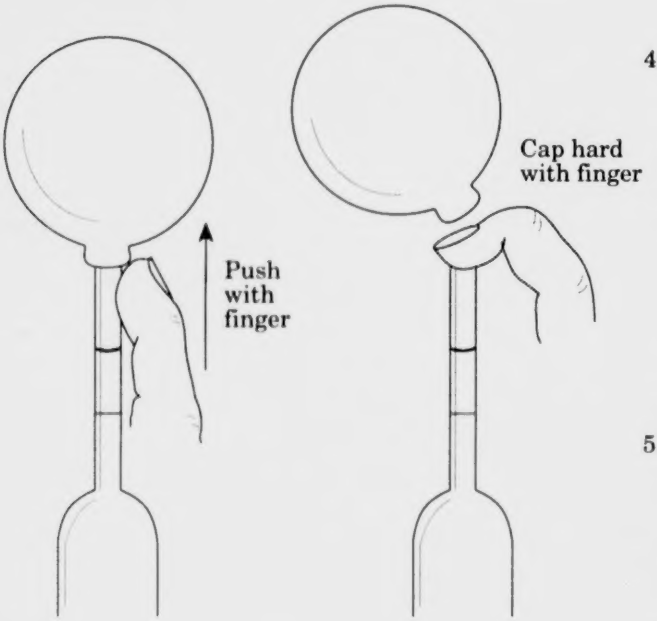


Figure 12. Removing the bulb.

4. Attach the bulb to the pipet stem to a depth of a few millimeters. (If the stem is inserted too deeply into the bulb, it will be difficult to separate the two.) Compress the bulb, then insert the tip of the pipet into the solution. Slowly release the pressure on the bulb allowing the solution to be drawn into the pipet. Continue releasing the bulb pressure until the liquid level is about halfway between the volume mark and the bulb. Students tend to compress the bulb after inserting the tip into the solution. This can lead to contamination and splashing. Stress using a "Blow no bubbles" technique!

5. Keeping the pressure on the bulb constant, push the bulb off the stem with the forefinger of the other hand. In the same action, quickly clamp the forefinger over the stem end. This will hold the liquid level above the volume mark (Figure 12).

6. Wipe the pipet tip with a clean towel. Release upper finger pressure ever so slightly allowing the solution to drain **just to** the upper volume mark. This can be done by slightly rotating the finger on the stem or by slowly relaxing the tension in the finger. The fleshy part of the forefinger gives the most flexibility and control. Use of the thumb to control pipet liquid levels is clumsy and should be discouraged. The draining solution can be caught in the beaker holding the remainder of the solution. Touch the tip of the pipet to the beaker to remove any partially formed drops. If the liquid level falls below the volume mark or if any air bubbles are present in the pipet, drain the solution back into the original beaker and start again.

7. Hold the pipet tip against the inside surface of the flask that is to receive the measured volume. Remove the finger from the stem and allow the solution to drain (Figure 13) to the bottom volume mark of the pipet or completely in the case of transfer pipets that have no bottom mark. After draining, touch the pipet tip to the inside surface of the flask. This will remove the last required volume. There will be a small volume of solution still in the pipet tip of a transfer pipet. **Do not** blow out this remaining solution. The pipet has been calibrated to leave this volume undelivered.

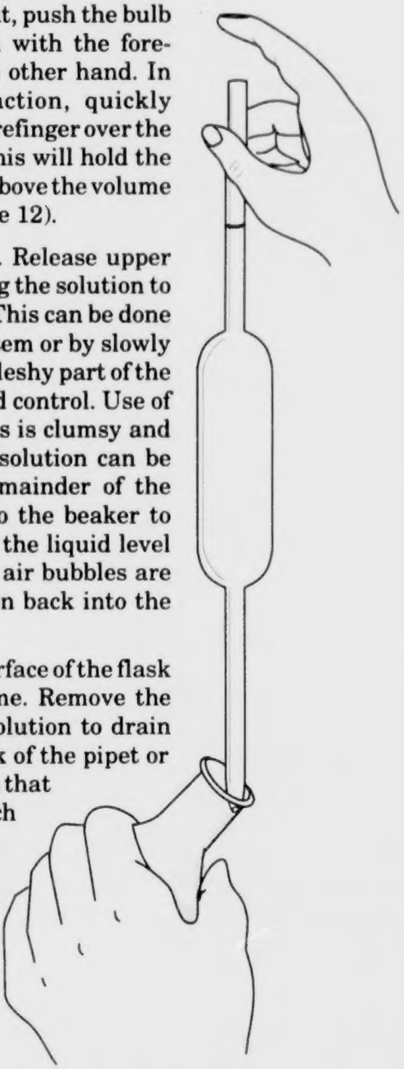


Figure 13. Draining a pipet.

8. If the pipet is used to measure a solution of the same substance but at a different concentration, rinse the pipet twice with the new solution before measuring. If a solution of a different substance is to be measured, rinse the pipet three times with distilled water then two times with the new solution. Discard each rinse. When finished with pipet work, rinse the pipet three to five times with distilled water as described above in Step 1 and place upside down to drain.

Volumetric Flasks

Volumetric flasks hold measured volumes and are used to prepare solutions of known molar concentrations. The flasks are precise to as much as ± 0.02 %. Flasks marked as Class A have the highest precision and more expensive. Class B flasks are not marked at all, but are quite acceptable for high school use. One must be concerned with contamination as well as precision in using the flasks, so cleanliness is imperative!

In the following instructions, it is assumed that distilled water is the solvent. *If this is not the case, use the desired solvent rather than distilled water.*

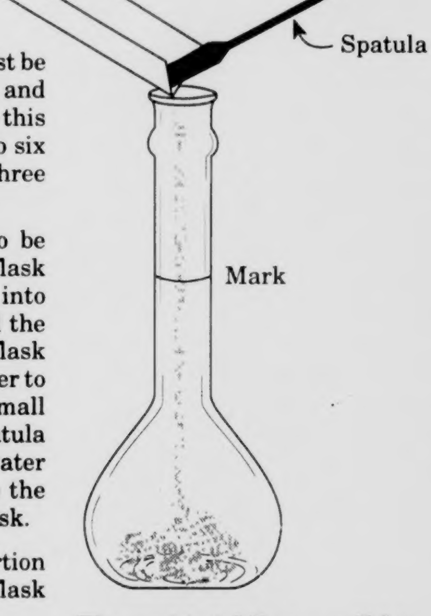


Figure 14. Adding a solid to a volumetric flask.

1. Rinse the flask three times with distilled water. This can be done by putting a small amount (about 1/5 the flask volume) of water in the flask and then rotating the flask until the entire "fat" portion and stem have been repeatedly coated.

   If you observe an uneven water coating, the flask is dirty and must be cleaned. Put a small amount of soap solution in the flask. Shake and rotate the flask until the entire inside surface is coated. Repeat this treatment twice with fresh soap solution. Rinse the flask four to six times with tap water to remove all traces of soap. Then rinse three times with distilled water.

2. Quantitatively (precisely) transfer the solute (the material to be dissolved) into the flask. If the solute is liquid, measure it into the flask using a pipet or a buret. If it is a solid, accurately weigh the solute into a small beaker or onto a creased piece of weighing paper. Hold the beaker or paper against the rim of the flask opening (Figure 14). Tip the beaker or paper to a slight angle and use a spatula to scoop small amounts of solid into the flask. Rinse the spatula and the beaker or paper with distilled water catching *all* the rinse in the flask. Rinse the spatula and paper twice more into the flask.

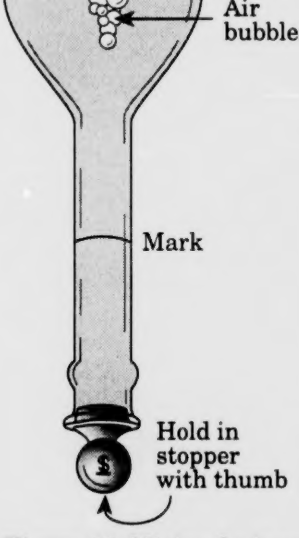


Figure 15. Mixing by inversion.

3. Add distilled water until the "fat" portion of the flask is about 4/5 full. Swirl the flask to thoroughly dissolve the solute.
4. Carefully add distilled water until the liquid level is **just at** the volume mark. Stopper the flask. Mix the solution by inverting the flask (Figure 15). This is done by quickly turning the flask upside down allowing the air gap to move through the solution. Mix by inversion allowing the air bubble to traverse the flask completely. Repeat several times to assure the solution is homogeneous. Be sure to hold the stopper firmly in place. (Mixing by shaking simply does not work well with volumetric flasks.)

After mixing and removing the stopper, you will often notice that the liquid level is slightly below the volume mark. This "decrease" in volume may occur due to solvation of the solute. This is simply because some of the solution is now coating the upper stem and stopper. Do not be concerned and **do not** add more distilled water.

5. Transfer the solution to an appropriate storage container. Volumetric flasks are easily knocked over and too expensive to use for solution storage.

Notes on Volumetric Flask Care

1. Never dry volumetric flasks in a drying oven. This treatment can ruin calibration of a flask. Flasks need not be dry before use to prepare aqueous solutions.
2. Store clean volumetric flasks about half full of distilled water to provide ballast so they cannot be knocked over easily.
3. Keep the standard taper stopper in the neck of the flask but place a small piece of weighing paper between the neck and stopper to prevent "freezing" of the stopper.
4. Never prepare concentrated basic solutions in Class A flasks. Caustic solutions etch glass and "freeze" the stopper.

Burets

Burets are calibrated to measure the volume of liquid delivered and are precise to ± 0.02 mL. Burets are used whenever volumes are to be delivered or more commonly, in titration analysis experiments.

The purpose of a titration is to determine the precise volume of a solution needed for a reaction. This is usually done by using the buret to add small amounts of the solution to the reaction flask until an indicator changes color (the end point). The end point may also be detected by use of pH meters and other instruments. The volume of solution added is determined by noting the buret volume readings before and after these additions. The solution in the buret is called the **titrant**. The point in the experiment where the indicator permanently changes to the desired color is called the **end point**.

Preparing the Buret

In order not to contaminate or dilute the solution being measured (the titrant), the buret (Figure 16) must be cleaned and the inside surface coated with the titrant before use.

1. Rinse the buret three times with distilled water. This can be done by filling the buret 1/5 full with distilled water and then tipping and rotating it to coat the entire inside surface with the water. Next, hold the buret upright and turn the stopcock allowing the water to drain into a "refuse" beaker. Note the coating of water on the buret inner surface.
2. If the water coating is uneven, showing breaks or droplets, the buret is dirty and should be cleaned with soap solution. Add some soap solution and clean the entire barrel using a buret brush. If brushes are not available, rinse the buret three times with soap solution using the procedure outlined in Step 1. Be sure to drain some soap solution through the stopcock and tip.

   Rinse buret barrel, stopcock, and tip about five times with tap water to remove all traces of soap. Then rinse these parts three times with distilled water.
3. Using the procedure in Step 1, rinse the buret barrel, stopcock, and tip at least twice with the titrant solution. This will replace the water coating with a coating of titrant.

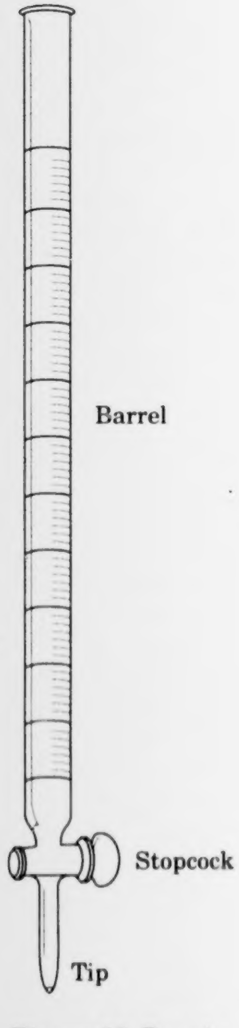


Figure 16. Buret.

Titration Procedure

Follow these steps in performing a titration:

1. Secure the buret to a stand using a buret clamp. Fill the buret with titrant to a point above the zero volume mark (a small funnel is helpful). Place the refuse beaker under the buret and open the stopcock full force allowing the titrant to drain a bit below the zero mark. This full force drainage is meant to eliminate any air gaps in the stopcock and tip. Check the tip and barrel for air gaps and bubbles. Dislodge any air bubbles by gently tapping while draining the buret full force.
2. Record the volume reading on the buret. (The initial volume reading need not be at zero.) Remove any drops on the tip by touching the tip to the side of the refuse beaker.
3. Place the solution flask to be titrated under the buret. Be sure the indicator has been added to the flask. By carefully turning the stopcock, add *small amounts* of titrant to the flask. Swirl the flask after each addition to completely mix the two solutions (Figure 17).
4. Continue additions until one drop of titrant causes a permanent (30 s) color change in the swirled flask. This is the end point of the titration. Record the volume reading of the buret and calculate the total volume of titrant solution that has been added.

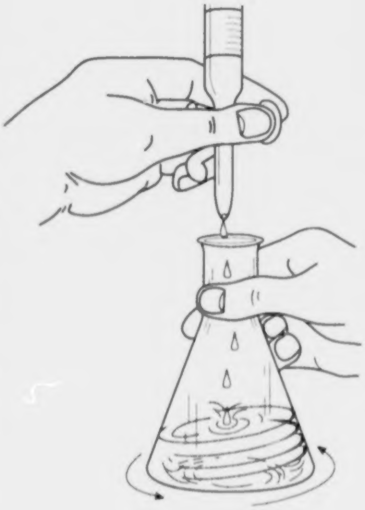


Figure 17. Swirl flask after each addition.

Pointers on Titrating

The following pointers can improve titration accuracy:

1. Make buret readings as precisely as possible. Determine the readings to within a hundredth of a milliliter by estimating the distance between volume markings. Wait 15 s after adding titrant before making a reading. This allows the solution on the sides of the barrel to finish draining. If you find it difficult to see the exact bottom of the meniscus, place a darkened piece of paper behind the buret about 1 cm below the meniscus. The meniscus will "reflect" this darkening and become more distinct (see Figure 18).
2. Locate the end point within one drop (generally 1/20 mL). Adding one drop at a time through the entire titration takes much time and is boring. Alleviate this as follows. Begin by adding about 1/2 to 1 mL at a time and watch the coloring where the titrant hits the solution. If the color change dispels by itself, the end point is not close. If the flask must be swirled to dispel the color change, reduce the size of the addition. When two or more swirls are needed to dispel the color change, the end point is close and single drop additions should be made. Alternatively, if there is a reasonable idea where the end point should be, add 80% of the needed titrant in one delivery and then begin adding dropwise. Practise manipulating the stopcock with one hand while swirling the flask with the other hand. This practice hastens a titration.
3. Place a white sheet of paper under the flask. The color changes can be seen more clearly.
4. Near the end of the titration, rinse down the inside surface of the flask with a little distilled water from a wash bottle to return all substances to the body of the solution.

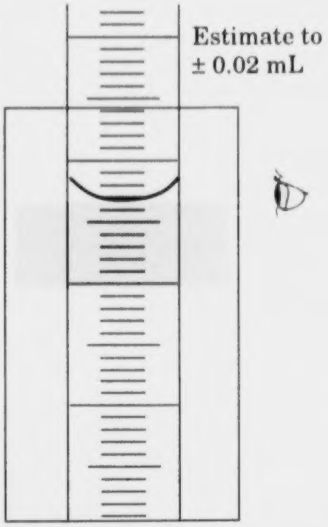


Figure 18. Making the meniscus clearer.

5. Do not trust the results of a single titration. If possible, repeat the titration of a particular solution several times to check on reproducibility. If you titrate several times in succession, start with the smallest amount of analyte (substance being analyzed) you will use; succeeding samples will require greater volumes of titrant so you may quickly add the previous volume of titrant to the new sample and continue slowly from there.
6. If the end point is passed (overshot) by adding too much titrant, add dropwise more of the analyte. This is called *backtitrating* and can be used intentionally to speed the titration since the end point need not be approached so cautiously. This procedure should be used only when the analyte is a liquid or in solution.

Nomenclature Guide

TIPS ON ORDERING CHEMICALS

When ordering chemicals, note the chemical grade and purity of each item. Reagent ACS and reagent grade chemicals are of the highest purity and hence the most expensive. Less expensive, and of less purity, are laboratory grade and technical grade. Both of these are suitable for most high school laboratory work. Often, the less pure (and less costly) technical grade is suitable. Unless students are involved in exacting analytical work, they should not need high purity chemicals. Chemicals of lower grades are often available at drug, hardware, and automotive stores. It may save both money and ordering time to investigate these local sources. A cost comparison of various grades of sodium chloride and ethyl alcohol is provided in Figure 19.

Substance	Grade	Amount	Cost (Dollars)
Sodium Chloride	99.5% minimum	500 g	5.60
	Laboratory grade	500 g	4.10
	Technical	1 kg	6.05
	Grocery store	454 g	.40
Ethyl Alcohol	Reagent, anhydrous specially denatured	500 mL	8.50
	Denatured, anhydrous	500 mL	5.25
	95%, Denatured	500 mL	5.25
	70% in water	500 mL	6.10
	Burner fuel	4 L	19.05
	190 Proof (grain alcohol) liquor store	750 mL	11.00

Figure 19. Comparison of chemical grades with respective prices.

STOCKROOM MANAGEMENT

It is absolutely imperative that the chemistry stockroom be a secured and safe area. Most school districts have safety or risk management officers who can assist you in planning chemical storage to avoid possible interaction of noncompatible substances. It is best to order smaller quantities, which can be used while fresh, and to regularly purge the stockroom of out-of-date substances. Many suppliers provide storage suggestions upon request. Each chemical ordered will be accompanied by an MSDS (Material Safety Data Sheet), which should be read first and then filed for ready access. A stockroom should be *Off Limits* to students and teachers who have not been trained specifically in its use. (For details on storage and safety, see the *Safety* module.)

IDENTIFYING SUBSTANCES: NAMES *VS.* STRUCTURES AND FORMULAS

Besides the purity of a substance you wish to order, there is also the problem of identifying it. This is because chemical supply house catalogs often use a substance's common name, while handbooks give more systematic names. This is especially true for organic compounds.

The systematic names, adopted by the International Union of Pure and Applied Chemistry (IUPAC), allow one to derive the molecular structure from the name, and *vice-versa*. Details of this system are included in college chemistry texts.(See *References*) Suppliers often use old or *trivial* names or some system of their own design. Examples of the variety of names for a single substance are shown in Figure 20.

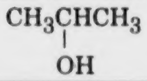
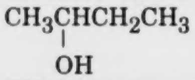
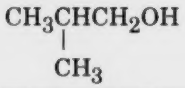
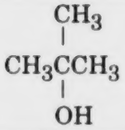
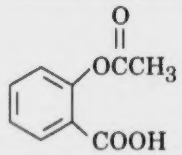
Molecular Formula	Structural Formula	Names
CH_4O	CH_3OH	Methyl alcohol, methanol, wood alcohol
C_2H_6O	CH_3CH_2OH	Ethyl alcohol, ethanol, grain alcohol
C_3H_8O	$CH_3CH_2CH_2OH$	*n*-propyl alcohol, propyl alcohol, 1-propanol
	 $CH_3\underset{\underset{OH}{\vert}}{C}HCH_3$	*i*-propyl alcohol, isopropyl alcohol, 2-propanol, rubbing alcohol
$C_4H_{10}O$	$CH_3CH_2CH_2CH_2OH$	*n*-butyl alcohol, butyl alcohol, 1-butanol
	 $CH_3\underset{\underset{OH}{\vert}}{C}HCH_2CH_3$	*s*-butyl alcohol, *sec*-butyl alcohol, 2-butanol
	 $CH_3\underset{\underset{CH_3}{\vert}}{C}HCH_2OH$	*i*-butyl alcohol, isobutyl alcohol, 2-methyl-1-propanol
	 $CH_3\overset{\overset{CH_3}{\vert}}{\underset{\underset{OH}{\vert}}{C}}CH_3$	*t*-butyl alcohol, *tert*-butyl alcohol, 2-methyl-2-propanol
$C_9H_8O_4$		Acetylsalicylic acid, aspirin, salicylic acid acetate

Figure 20. Substance names.

For inorganic compounds, the Stock system replaces *-ous* and *-ic* with Roman numerals. The ferrous ion, Fe^{2+}, is iron(II), while the ferric ion, Fe^{3+}, is iron(III). The Roman numeral thus indicates the ion's charge. The Stock name for CuCl is copper(I) chloride (old name is cuprous chloride) and for $CuCl_2$, copper(II) chloride (old name is cupric chloride). When the oxidation number of an ion is written using Arabic numerals, such as Fe^{2+} for iron(II), the notation is known as the Ewens-Bassett system. Stock and Ewens-Bassett notation is used interchangeably throughout *SourceBook*. Handbooks, as well as most supply catalogs, now use this system of naming inorganic salts. Many chemical catalogues have an extensive and cross-referenced chemicals list. They contain descriptions of physical properties and are useful and free reference books. Two excellent sources of information and reagents for the high school laboratory are the *Aldrich Catalog of Fine Chemicals* and the *Flinn Chemical Catalog and Reference Manual*.

Unit Conversions: The Factor Approach

Most answers to chemistry problems have two parts: a number followed by a unit or units. Exceptions are logarithms and ratios. Proper use of units is an aid when solving problems. The technique is to treat units as algebraic quantities and carry them through all calculations. To use a simple example: Find the time, in hours, for the 400 km trip from City A to City B, if travelling at an average speed of 80 kilometers per hour for the trip. First, look at the given units, km/h and km. Multiply h/km by km because the km in the first term will then divide the km in the second, leaving an answer in hours.

$$\frac{\text{h}}{\text{km}} \times \text{km} = \text{h}$$

To solve the problem, place the numbers in front of these units: $\frac{1\ \text{h}}{80\ \text{km}} \times 400\ \text{km} = 5\ \text{h}$

Note that km factors (cancels) in the same way that the number 2 factors (cancels) in the problem $\frac{2 \times 3}{5 \times 2} = \frac{3}{5}$

Such an approach to problem solving is known as *dimensional analysis* or *unit analysis.*

It is critical to realize the *same* quantity can be represented by different numbers. For example, the length of a 12-inch ruler, can be represented by *different* numbers: 12 inches and 30.5 cm. A useful exercise is simply to look at an English-metric ruler, a portion of which is reproduced in Figure 21. It is immediately obvious that the *same* length can be represented by *different* numbers because the *units* are different.

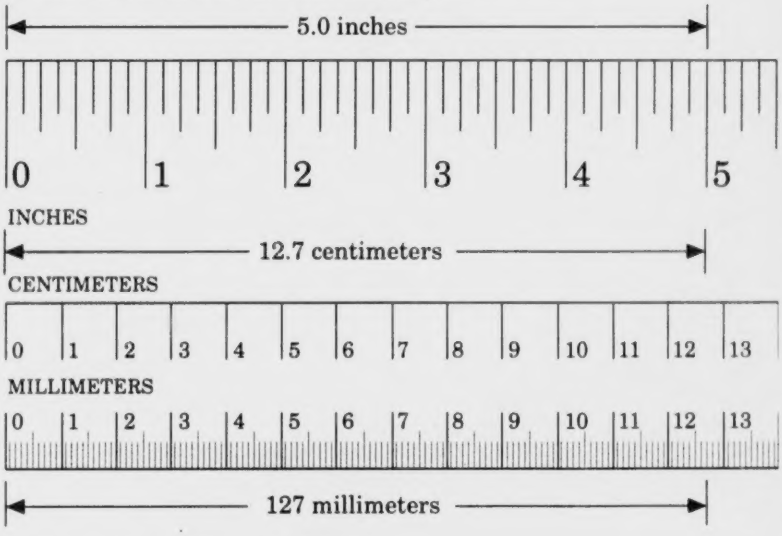


Figure 21. English-metric ruler.

Additional examples of the use of dimensional analysis are given below.

To calculate the number of drops of water in a lake, one might follow this process:

$$\frac{\text{Drops}}{\text{Lake}} = \frac{\text{Drops}}{\text{mL}} \times \frac{\text{mL}}{\text{L}} \times \frac{\text{L}}{\text{Lake}}$$

To find the number of liters in the lake, one could estimate:

$$\frac{\text{L}}{\text{Lake}} = \frac{\text{m}^2\,(\text{area})}{\text{Lake}} \times \frac{\text{m (depth)}}{1} \times \frac{10^6\text{cm}^3}{\text{m}^3} \times \frac{1\text{L}}{10\text{cm}^3}$$

Use of dimensional analysis in chemistry will be employed extensively in problem solving techniques below. Because many students do not see fractions as division, this equivalence should be emphasized.

Uncertainty in Measurement

ACCURACY *VS.* PRECISION

Accuracy is defined as the closeness of a measured quantity to its true or accepted value. Precision, in contrast, refers to the reproducibility of measured data. A useful way to help students distinguish between these often-confused terms is to use a target practice analogy. Here, a beginner and an expert each fire a rifle at a target. The appearance of the targets is shown in Figure 22. The novice shot at Target A and shows neither accuracy nor precision. The expert fired first at Target B, with a high degree of precision. All bullets hit the same area of the target. Clearly, the rifle's sight scope is out of alignment. Target C shows the expert's results after correcting the sighting scope. It indicates both high precision and good accuracy.

It should be apparent that an experimenter never knows the accuracy unless items that have known values are measured or some other measurement confirms the values. Precision is measurable, but it does not mean that we have been accurate. Furthermore, time is limited in most school settings, and students seldom repeat measurements sufficiently to determine precision.

We are left with the uncertainty of the result. This uncertainty arises from determinate and indeterminate errors. The former type of error includes miscalibrated instruments, nonideal behavior of substances (such as absorption of moisture by a sample during weighing), errors in personal judgement or from physiological restrictions (*e.g.*, poor vision). Indeterminate errors happen, and we never know why (an unknown cosmic ray shower's effect on electronics for example). These errors are the sum of the little things that could go wrong. Murphy's Law applies here.

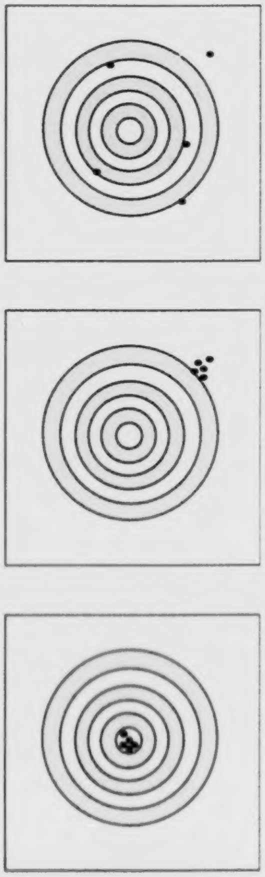


Figure 22. Accuracy vs. precision.

Strive to instill the concept that no measurement is perfect, and it is the scientist's responsibility to report honest results and communicate to the scientific community the quality of the experimental work. The following sections summarize frequently used methods to accomplish this goal.

RULES FOR UNCERTAINTY IN MEASURED QUANTITIES

Every time students make measurements in the laboratory they face the uncertainty that is part of any experimental science. Regardless of how carefully students work, the limitation of the measuring instrument itself means the measurement must be uncertain. Students should be encouraged to record not only their measurements, but the uncertainties in each measurement. Since those uncertainties are built into the instrument, students need to be aware of the uncertainties in each instrument used.

Several types of laboratory balances (Figures 5-7) are pictured in this module. Suppose a platform balance (Figure 23) has the limitation that it appears balanced over a 0.1 g interval. One student might read and record a mass as 61.2 g while another could report 61.3 or 61.1 g. Each measurement is known only to the nearest 0.1 g. The first reading should therefore be reported as 61.2 ± 0.1 g, where the ± 0.1 g notation means that, because of uncertainty, the actual mass could be as high as 61.3 g, or as low as 61.1 g. The measurement is said to have an uncertainty of ± 0.1 g. Since the platform balance cannot be more precise, this notation is the most honest and most informative way of reporting the measurement. It is honest because it admits there is uncertainty. It is informative because it reveals the extent of uncertainty.

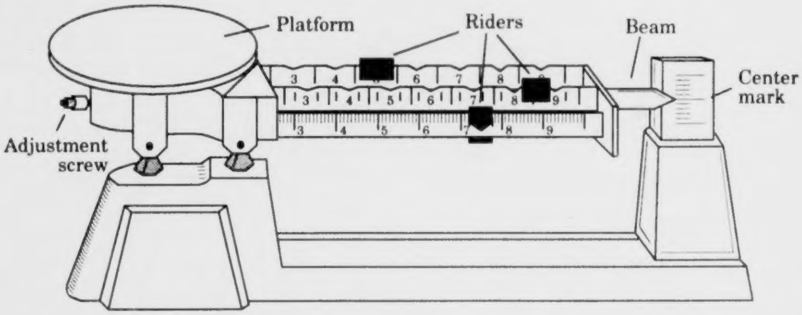


Figure 23. The platform balance.

A triple beam balance (Figure 7) appears in balance over only a 0.01 g interval. Its uncertainty is only ± 0.01 g. Using it, a student might report a mass of 61.18 ± 0.01 g. This indicates the actual mass could be as low as 61.17 g or as high as 61.19 g. The electronic balances have even less uncertainty, so mass measurements on them can be reported to an even greater precision.

In a similar way, volumetric measurements with graduated cylinders, pipets, and burets are also uncertain, so that measurements made with them must include the uncertainties in these measurements. *Note in each case that the last decimal place to be kept in a measurement is the uncertain place, and that the ± notation suggested above indicates how uncertain that number is.*

RULES FOR UNCERTAINTY IN DERIVED RESULTS

When students use uncertain quantities to obtain calculated results, their calculators often give them six, seven, or eight place answers. Remind them that such exact answers may not be justified. How do you best convince them? Many preliminary calculations involve addition or, more often, subtraction. If weighing paper or a bottle is used, the mass of the object is the mass of the object plus that of the empty paper or bottle, minus the mass of the empty paper or bottle. Common sense says that, since the first mass may be too high, the second too low, and we have no way of knowing which, the derived mass must have an uncertainty that is the *sum* of the directly measured uncertainties.

The same thing is often true when measuring volumes. Suppose, for example, that a student's buret liquid level was 5.20 mL at the start of a titration and 15.40 mL at the end point. If the buret's uncertainty was ± 0.02 mL, the first reading could have been as high as 5.22 or as low as 5.18 mL; the second reading, as high as 15.42 or as low as 15.38 mL. The maximum-minimum approach suggested below should convince the student that the above rule of adding uncertainties is valid. Here, the maximum difference would be 15.42 – 5.18 = 10.24 mL. The minimum, 15.38 – 5.22 = 10.16 mL. The directly calculated value, 15.40 – 5.20 = 10.20 mL, could therefore be as high as 10.24 and as low as 10.16, or inbetween. Thus, the best way to report this calculated result is 10.20 ± 0.04 mL. The ± 0.04 mL uncertainty is the sum of the 0.02 mL and 0.02 mL uncertainties in each originally measured volume. It is thus unnecessary to do a maximum-minimum analysis; simply use the rule of adding the original uncertainties. This rule also holds for subtraction.

The rules for uncertainties when multiplying or dividing uncertain quantities are less straightforward but are also amenable to the maximum-minimum approach. For example, a large, rectangular laboratory's dimensions are hurriedly measured as 20 m by 10 m with each dimension precise only to the nearest meter. Dimensions would be properly reported as 20 ± 1 m by 10 ± 1 m. Maintenance wishes to cover the floor with chemically resistant floor tiles and needs to know how many square meters to buy. What do you advise?

A maximum-minimum approach shows that the largest possible area would be 21 m x 11 m = 231 m^2, while the smallest is 19 m x 9 m = 171 m^2. If the dimensions were indeed 20 by 10 m, the area would be 200 m^2. Since exact dimensions are not known, the floor area could be some 30 m^2 more, or less, than this and would best be reported as 200 ± 30 m^2, approximately. Maintenance should buy just over 230 m^2 of tiles to be safe.

Clearly, ± 30 is *not* the sum of ± 1 and ± 1. What is the rule with uncertainties when multiplying uncertain quantities? Consider percent uncertainties. One part in 10 is 10%, one part in 20 is 5%. Thirty parts in 200 is 15%, so the approximate rule when multiplying is to add the percent uncertainties, 10 + 5, rather than the so-called absolute uncertainties, 1 + 1. To convert the 15% to an absolute uncertainty, simply take 15% of 200; that is, 0.15 x 200 = 30. The same approximate rules apply to division, which is mathematically similar to multiplication.

In summary, the rules just developed serve as guides for finding the uncertainty in any result calculated from uncertain data. While the precision of any such result can be determined by maximum-minimum analysis, the rules provide a shorter alternative. Thus:

1. A calculated result is reported to the same number of places as the data used to obtain it. The last place retained is the uncertain place. In the 200 m^2 area example, the tens place is uncertain.
2. When adding or subtracting uncertain quantities, add their absolute uncertainties.
3. When multiplying or dividing uncertain quantities, add their percent uncertainties.

If the percent uncertainties seem unwieldy for students, simply use rule (1). In any case, report all calculated results as precisely as justified. The *Thermochemistry* module provides more chemistry-related examples of uncertainties in calculated results.

SIGNIFICANT FIGURES AND SCIENTIFIC NOTATION

What are Significant Figures?

All figures that are certain in a measurement plus one uncertain figure are known as the *significant figures*. The last figure shows where uncertainty begins. If a thermometer indicates a melting point of 31.2 °C, and has an uncertainty of ± 0.2 °C, all figures in that 31.2 are significant, including the 0.2, which is uncertain. The notation 31.2 ± 0.2 °C indicates how uncertain it is.

To extend this to a calculation, suppose the floor area of another rectangular laboratory is 651 m^2 with an uncertainty of ± 30 m^2. The area would be reported as 651 ± 30 m^2, but that makes no sense. If the 5 in the tens place is uncertain, we cannot know the unit place 1. The area should be 650 ± 30 m^2 where only the 6 and 5 are significant. The zero is not significant. With floor dimensions of 21 and 31 m, the calculated area can only have two significant figures, since the original measurements (the 21 and the 31) each have only two significant figures.

Reporting the 650 m^2 area without the uncertainty leads to confusion about whether the zero is significant. Use of scientific notation removes this confusion since the area would be reported as 6.5×10^2 m^2. The 6.5 has only two significant figures. The nonsignificant zero *is* needed in the 650 notation to place the decimal and show that the answer is *not* 65. That zero is *not* needed with scientific notation.

Sometimes it is convenient to make a table more compact by omitting a repetitious exponential term. When an exponential term is listed as part of a column heading (Figure 24, Column A), it means the original value (Figure 24, Column B) was multiplied by the heading value prior to inclusion. Unfortunately, this is sometimes seen in reverse in publications (Figure 24, Column C). It is advisable to check the text for confirmation of the system in use.

A
10^3 x Moles
1.22
1.18
2.01

means

B
Moles
0.00122
0.00118
0.00201

not

C
Moles
1220
1180
2010

Figure 24. Exponents used in column headings.

Calculating Uncertainty by the Significant Figures Method

Recalling that significant figures are the result of measurement, their name is quite sensible. They are figures (numbers, digits) that have significance (meaning, value) in the context of the laboratory work. It is not significant to look at a clock and say the time is 2:15:26.0000. The second hand may point at 26, but it is moving and we have no chance at interpolation. It is nonsense to say more than 2:15:26. Similarly, a person's weight according to a bathroom scale could not be reported meaningfully to be 120.0178 lbs; most such scales have uncertainties of more than a pound. Stress that data must reflect a reasonable ability to measure it with the equipment provided.

We can simplify dealing with uncertainty by use of a few simple rules based on the presumptions that:

1. All laboratory readings are made to the limit of detection (minimum uncertainty) of the measuring device.
2. All measurements written down include only significant digits and identifiable indicators of magnitude—(place holding zeros or exponential notation).
3. The last significant figure in the measurement is admitted to be ±1 in its column. Figure 25 illustrates these rules. Reading the buret gives 51.29 mL. The meniscus is certainly between 51.2 and 51.3 but we are uncertain as to where. The best estimation is 0.9 of the distance between 51.2 and 51.3, but we could be off by 0.02 mL either way. Write the reading as 51.29, understanding it is ±2 in the hundredths column or 51.29 ± 0.02 mL.

In reading a digital balance the display indicates 127.395 g. Assuming an uncertainty of ± 0.001, notice that the balance switches the last digit from 5 to 4 to 6 as it stabilizes. This indicates that it, too, is uncertain but chooses the 5 as the best value.

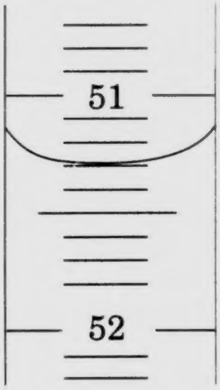


Figure 25. Buret meniscus (note curved surface).

Students always find formal study of uncertainties difficult and seemingly foreign. This anecdote points out that they do have an inherent understanding of the topic and that it makes "common sense" to think about significant figures.

> I once stopped at a roadside "tourist trap" in Arizona that advertised a display of dinosaur fossils. While viewing them I inquired of the attendant as to their age. He assured me they were 150,000,006 years old. I was really impressed with the quality of the geologic dating and asked how it was done. He said a university paleontologist had come down, looked at them and said they were 150 million years old. Since that was six years ago, he knew their age exactly.

Students always laugh at the dinosaur story. They know it does not work that way. The *exact* 6 years are not addable to the *crude* 150,000,000 years. It will be 10 million years before those bones will be any older, unless someone improves the significance of the age by measuring some characteristic of them that supports adding significant figures to the original age.

Keep in mind:

The final calculation is no better than the poorest measurement used to make it!

Rules for Use of Significant Figures in Data Processing

1. All measured nonzero digits are significant.
2. Zero may be significant. It will be if it is not a simple place holder (which would disappear in scientific notation). Examples are:
   a. Between nonzero digits: 303 g has three significant figures; 300 °C has only one.
   b. Following a nonzero digit and a decimal point: 300.00 g has five, 0.46000 g has five.
   c. In the following numbers, zero is not significant: 910 mL, 0.91 °C, and 0.091 g.
3. When adding or subtracting, the final answer may contain only those columns that were filled with significant figures.

Example:	198.7 sec	150,000,000 y	4.0145 km
	+ 17.43 sec	+ 6 y	–2.12 km
	215.1 sec	150,000,000 y	1.89 km

4. When multiplying, count the number of significant figures in each multiplier and round the final answer to the nearest number containing the same number of significant figures as the multiplier that contained the fewest. The final answer may contain nonsignificant zeros as place holders or be written in scientific notation without the leading or following zeros. Division and other nonadditive processes follow the same rule.

Example: 3 km/hr x 6 hr = 20 km 51.278 g/76.4 L = 0.0671 g/L
21.70 g/mol x 20.2 mol = 438 g $(2.3\ \text{cm})^2 = 5.3\ \text{cm}^2$

Note that all examples shown included units of measure. This was done to emphasize that significant figures are used to deal with laboratory values. Emphasize this to students. They have worked with integers in mathematics classes for many years. They find statistical treatment of data quite foreign. Stressing that laboratory numbers are different from "math class" numbers will help them adjust. Point out, also, that counting is statistically exact and such numbers as "12 eggs in a dozen" have no uncertainty and do not limit the certainty of a calculated result.

Example: 500 sheets/ream x 4 reams = 2000 sheets (no uncertainty)

Avogadro's number will never be counted. It has been experimentally determined to at least seven significant figures. There is of course an exact integral number of particles in a mole. It is just that we are uncertain of that number. In this small way, a mole is not just like a dozen.

Be careful not to overemphasize uncertainty early in the year. It tends to turn off students. They will appreciate it in time if you frequently remind them of the rules, point out sources of error during demonstrations and adhere to the rules scrupulously in examples and examination questions. It is an uphill battle, especially since most texts are quite inconsistent in observing good treatment.

Consider using the following exercise on precision, accuracy and uncertainty to convince students of these topics' importance.

> Use a reducing/enlarging copy machine to reproduce laboratory rulers (graduated in mm) at sizes slightly larger and smaller than true size as well as some that are true size. Be sure they are centered on the paper and not noticeably different.
>
> Place the rulers at laboratory stations in random fashion. Provide 3"x 5" cards that students see come from a new package.
>
> Have students, working in groups of 3-5 at each ruler, individually measure the length and width of the card privately and then combine their results. Have them determine their group mean and range (±).
>
> Reassemble the class, let each group report the size of the card and discuss the results. Be sure to point out that good experiments require good reference standards.
>
> Return the students to the laboratory with legitimate rulers and have them determine the length and diameter of a test-tube, repeat their calculations and return to class for a further discussion.

Problem-Solving Techniques

MOLAR MASSES

Atomic and molecular weights, more properly called molar masses, are routinely presented as relative—as related to some arbitrary standard. The currently accepted standard is the carbon-12 isotope taken as precisely 12 atomic mass units (symbol u but also frequently seen as amu). It may be easier for students to relate to a hydrogen atom as approximately 1 amu. Then, like the tenth grader being twice as heavy as a second grader, they can imagine a helium atom four times as heavy as a hydrogen atom, an oxygen atom as 16 times as heavy, and so on. Furthermore, in this era of biotechnology, students might agree that 100 clones of the big tenth grader must be twice as heavy as 100 clones of little second grader.

Once students accept the above argument, it follows that a dozen helium atoms would be four times as massive as a dozen hydrogen atoms. A dozen oxygen atoms would be 16 times as massive as a dozen hydrogen atoms. But what holds for the consumer's dozen must also hold for the chemist's "dozen"—the mole. A mole of helium atoms must be four times, and a mole of oxygen atoms 16 times, as massive as a mole of hydrogen atoms. This brings us to the essential point that students recognize the difference between molar mass and atomic or molecular mass—the mass of an individual atom or molecule.

THE MOLE—A CHEMIST'S "DOZEN"

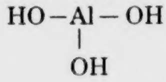
The molar mass of water, H_2O, is 18 g and because of the way the mole is defined, the mass of a single water molecule is 18 u. A single water molecule is some 18 times more massive than a single hydrogen atom, just as a mole of water molecules is some 18 times more massive than a mole of hydrogen atoms. The numbers, 18 u per molecule and 18 g per mol may be the same, but the *units are not*. Emphasize by asking students to use the *same* unit (grams) for both. The 18 g of water contain Avogadro's number, approximately 6 x 10^{23} water molecules. A single water molecule's mass, in grams, must then be:

$$\frac{18\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O} \times \frac{1\ \text{mol}\ H_2O}{6 \times 10^{23}\ \text{molecules}\ H_2O} = \frac{3 \times 10^{-23}\ \text{g}\ H_2O}{1\ \text{molecule}\ H_2O}$$

Written out, this becomes 0.000 000 000 000 000 000 000 03 g! Hardly the same mass as the 18 g found in one mole of those molecules!

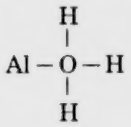
FORMULAS AND EQUATIONS

For beginning students, the language of chemistry may be as confusing as any foreign tongue. The precise nature of chemistry's language is well illustrated by the very different meanings of Co and CO. The former represents one atom of cobalt, the latter, one molecule of carbon monoxide. That NA, AL, MG are incorrect (properly Na, Al, Mg) may be difficult but essential for students to learn. When illustrating the proper use of subscripts, it is useful to diagram some formulas in terms of individual atoms. Here $Al(OH)_3$ represents an Al linked to three OH groups:



$HO-\underset{\underset{OH}{|}}{Al}-OH$

While $AlOH_3$ is incorrect because it represents:



$Al-\underset{\underset{H}{|}}{\overset{\overset{H}{|}}{O}}-H$

In contrast, $AlCl_3$ needs no parentheses, since Cl represents just one chlorine atom. Much practice will prove useful here.

Two accompanying activities can aid students with equation balancing. One, suggested in the *Simple Chemical Reactions* module section on student activities, uses different kinds of paper clips. Alternatively, use differently colored gumdrops joined with toothpicks. As shown in Figure 26, gumdrops represent atoms rearranging to form different configurations of gumdrops, that is, different molecules.

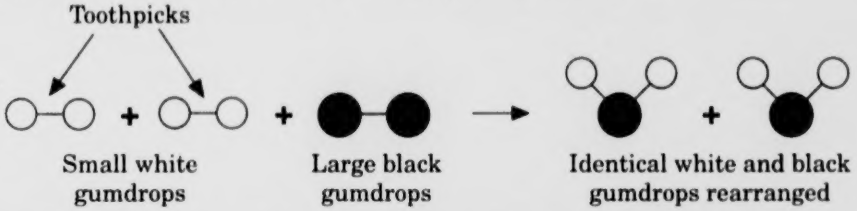


Figure 26. Rearranging atoms to form different molecules.

Note that while gumdrops are conserved (four white and two black on each side of the arrow), the number of configurations may not be. Here, all gumdrops present in the three arrangements to the left of the arrow are contained in the two configurations to the right. In the reaction shown in Figure 26, atoms are conserved, but not molecules. The rearrangement is analogous to the equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

A second activity, also suggested in the *Simple Chemical Reactions* module, Demonstration 3, involves mixing dilute colorless solutions of lead nitrate [$Pb(NO_3)_2$] and potassium iodide (KI) to produce a yellow precipitate, lead iodide, (PbI_2). If this is done on a balance pan, mass is observed *not* to change throughout this otherwise dramatic chemical change. This reaction is illustrated by the equation:

$$\begin{array}{lcl} \text{Pb-NO}_3 + \text{K-I} & \rightarrow & \text{Pb-I} + \text{K-NO}_3 \\ \quad\ \text{NO}_3 \quad \text{K-I} & & \quad\ \text{I} \quad \text{K-NO}_3 \end{array}$$

where taking inventory reveals one Pb, two K's, two I's, and two (NO_3)'s on *each* side of the arrow. You can convince students the equation is balanced by the technique in Figure 27. Use identical sets of alphabet blocks and set them up as shown in Figure 27.

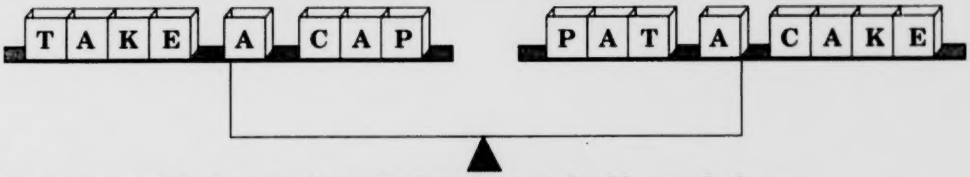


Figure 27. Alphabet blocks on both pans of a double pan balance.

Point out that each pan has the same eight blocks, and hence the same mass is on each balance pan. Switching some blocks gives very different messages. Here the letters represent atoms, and the words they form represent different molecules. A further point, emphasized by showing a *large* dictionary, is the many words that just 26 letters can form is analogous to the many compounds that the 104 known elements can form.

EQUATION STOICHIOMETRY

Students must understand what a balanced equation does and does not reveal about the reaction it represents. It tells neither how fast the reaction proceeds nor how *far* it proceeds. Many reactions do *not* go to completion, such as:

$$N_2 + 3H_2 \rightarrow 2NH_3$$

An equation does not indicate how much of each reactant is present, so it cannot predict how much of each reactant may remain after reaction. A balanced equation *does* indicate the relative number of each reactant species that *react*, and the relative number of each product species that will form. The formation of water from hydrogen and oxygen is represented by the reaction equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

Diagrammatic Representation

This equation is represented by each diagram in Figure 28.

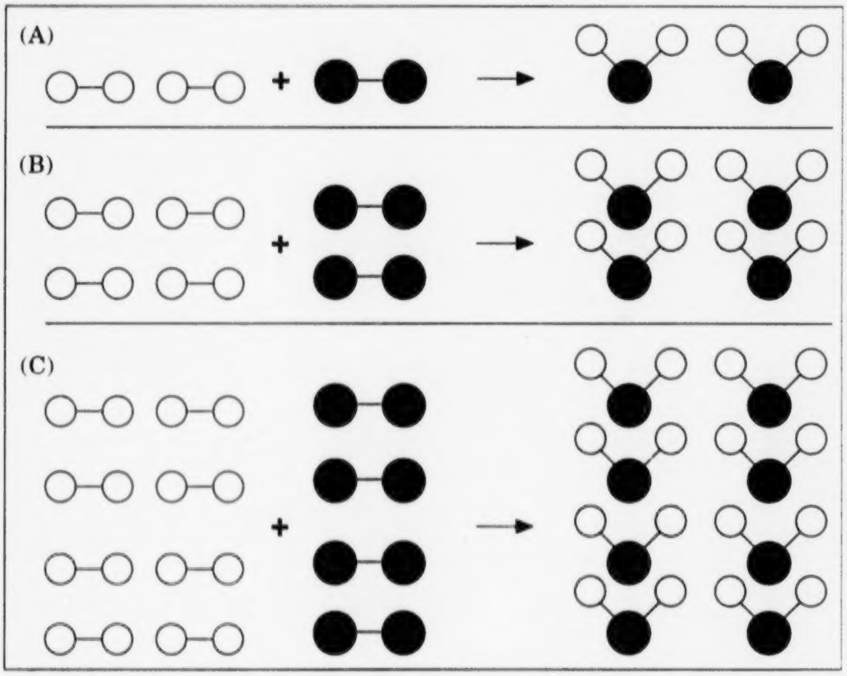


Figure 28. Various meanings of the equation $2H_2 + O_2 \rightarrow 2H_2O$.

In the figure, black circles represent oxygen atoms and small white circles, the hydrogen atoms. Note each of the diagrams A, B, and C show a two-to-one ratio of H_2 molecules to O_2 molecules. This same equation also represents the reaction between two dozen H_2 and one dozen O_2 molecules, or between two *moles* of H_2 and one *mole* of O_2 molecules. Equations can always be read as number of *molecules* or number of *moles* that react and form. For the formation of water reaction, the equation could also represent 1 mole of H_2 and 1/2 mole of O_2 forming 1 mole of H_2O since 1 and 1/2 are in a 2-to-1 ratio.

The essential information conveyed by the formation of water reaction equation is that each *one* O_2 molecule can react with *two* H_2 molecules, even if one of these species is present in excess at the start. *If* an excess of one reactant is present, that excess simply remains unreacted after the reaction is over. Thus two H_2 molecules can *react* with only one O_2 molecule, regardless of how many O_2 molecules may be present initially. The diagrams in Figure 29 help to remind students of this and should be carefully studied.

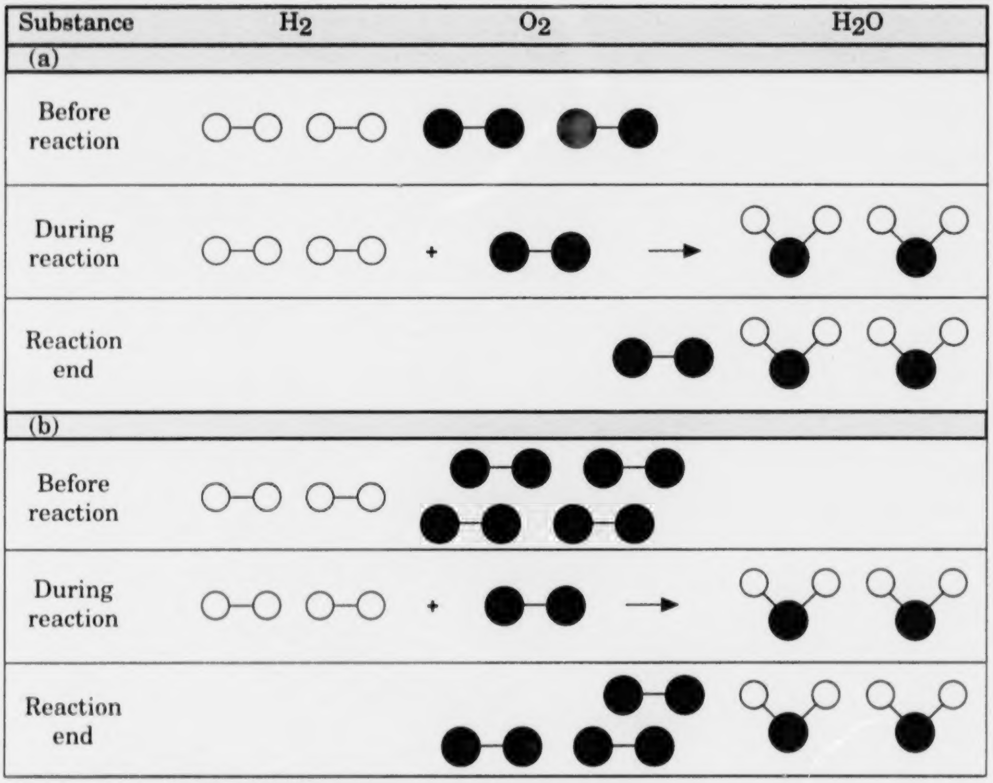


Figure 29. Excess of one reagent.

Note in Figure 29, only one of the two O_2 molecules initially present can *react*, so the other remains unreacted at the end, along with the two H_2O molecules produced. In Figure 29b, three of the initial four O_2 molecules remain unreacted. If students can master this "bookkeeping" method early on, equilibrium problems encountered later on in chemistry become easier since the "left at end" amounts are what remain at equilibrium. Once the meaning of a balanced equation is understood, all problems based on equations become workable. Consider several, based on the equation:

$$2H_2 + O_2 \rightarrow 2H_2O$$

This equation tells us that 2 moles of H_2 react with 1 mole of O_2 to form 2 moles of H_2O. Thus if 4 moles of H_2 react in *excess* O_2, only 2 moles of the O_2 will react, forming 4 moles of H_2O. Mole amounts of reactants and products are read directly from the equation coefficients, one must remember these represent only the simplest whole number ratios.

Solving Numerical Problems with Masses

If we are given *masses* of reactants, these must be converted in each case to moles, since equations only give mole ratios. A review of the unit factor approach may be helpful. Since molar mass of any substance is the same *number* of grams as its molecular mass in u, appropriate unit factors can be derived from the substance's formula and a table of atomic weights. Consider a problem in which we are given 10 grams of H_2 and 16 grams of O_2 as reactants. How many moles of H_2O will result from the reaction? Since molar masses of H_2 and O_2 are $1 + 1 = 2$, and $16 + 16 = 32$, respectively, we convert the given masses to moles:

$$10\ g\ H_2 \times 1\ mol/\ 2\ g = 5.0\ mole\ H_2;\ 16\ g\ O_2 \times 1\ mol/\ 32\ g = 0.5\ mole\ O_2$$

Since 0.5 mole O_2 can *react* with only 1.0 mole H_2 (twice as much), only 1.0 mole of H_2O will form. We could now convert the *moles* of H_2O formed into its mass, in grams. Since 1 mole H_2O = 1+1+16 = 18 g, 1.0 mole H_2O x 18 g/mol = 18 g of H_2O that would form. This, of course, is the sum of the actual reacting masses.

Consider the reaction between sulfur and oxygen gas. Sulfur, S_8, burns in oxygen gas, O_2, to form sulfur dioxide, SO_2. If 512 grams sulfur burn in just 320 grams oxygen gas, find the mass, in grams, of sulfur dioxide that can form?

The **general method** for solving such a problem involves these steps:

1. Translate the words into a skeleton equation: $S_8 + O_2 \rightarrow SO_2$
2. Balance the equation: $S_8 + 8O_2 \rightarrow 8SO_2$
3. Enter the "given" and "to find" information stated in the problems *above* the appropriate formula in the balanced equation, thus:

$$\begin{array}{cccc} 512\ g & 320\ g & & x\ g \\ S_8 + & 8O_2 & \rightarrow & 8SO_2 \end{array}$$

4. Calculate the molar masses for each substance in the equation, placing these *below* each formula represented. From these formulas and atomic masses, the molar masses are 256 g mol^{-1} for the S_8, 32 g mol^{-1} for the O_2, and 64 g mol^{-1} for the SO_2.

$$\begin{array}{ccccc} 512\ g & & 320\ g & & x\ g \\ S_8 & + & 8O_2 & \rightarrow & 8SO_2 \\ 256\ g\ mol^{-1} & & 32\ g\ mol^{-1} & & 64\ g\ mol^{-1} \end{array}$$

5. Find the number of moles of the given reactants. This involves converting the 512 g S_8 and 320 g O_2 to moles. Hence,

$$512\ g\ S_8 \times 1\ mol\ S_8/\ 256\ g\ S_8 = 2.00\ mol\ S_8$$
$$320\ g\ O_2 \times 1\ mol\ O_2/\ 32\ g\ O_2 = 10.0\ mol\ O_2$$

6. Read the coefficients in the balanced equation for the proper mole ratios.

$$1\ mol\ S_8 + 8\ mol\ O_2 \rightarrow 8\ mol\ SO_2$$

So, $$2\ mol\ S_8 + 16\ mol\ O_2 \rightarrow 16\ mol\ SO_2$$

From these relationships we see that 2 moles S_8 require 16 moles of O_2 to burn completely, but only 10 moles of O_2 are available. Therefore, O_2 is the "limiting reactant" and controls the amount of product that can form and limits the amount of sulfur that will be used. We may establish the moles of product from the equation ratio.

$$\frac{8SO_2}{8O_2} = \frac{x}{10.0\ mol\ O_2}$$
$$x = 10.0\ mol\ SO_2$$

If given definite amounts of *both* reactants, as in this problem, use a "bookkeeping" record of initial moles, moles reacting, and moles remaining at the end of the reaction. This approach will be particularly useful in equilibrium systems studied later.

The balanced equation:	S_8 +	$8O_2$ →	$8SO_2$
Moles initially:	2.00	10.0	0
Moles reacting:	1.25 +	10.0	10.0
Moles remaining at end:	0.75	0	10.0

Had there been excess O_2, then all 2.00 mol of S_8 would react, giving 16 mol of SO_2, which could then be converted into grams, as above.

7. Convert the moles of the "to find" substance, here the SO_2, into the requested unit, grams:

   10.0 mol SO_2 x 64 g/mol = 640 g SO_2, the desired answer

8. A compact representation of this process helps many students:
   a. Write and balance equation.
   b. Convert given mass (or volume) to moles.
   c. Use equation ratio to find moles of unknown.
   d. Convert moles of unknown to mass (or volume).

Numerical Problems with Gases

When students learn that one mole of any gas at standard conditions (0 °C and 1 atmosphere) occupies 22.4 liters in the *Gases* module, it is simple to calculate the *volume* (in liters) of any gaseous reactant or product. For example, consider the reaction between sulfur and oxygen detailed above. Since, for *any* gas at standard conditions (STP), 1 mole = 22.4 liters, this equality can be made into a unit conversion factor. Specifically, and again referring to the sulfur-oxygen reaction, both O_2 and SO_2 are gases. Hence the 10 moles of O_2 must occupy 10 mol O_2 x 22.4 L/mol = 224 L O_2 at STP, and the 10 moles of SO_2 produced must have this same volume, by the same conversion calculation.

Alternative Stoichiometry Methods

Students often find a ***"road map"*** (Figure 30) a useful organizer for solving problems. Two general "maps" are given below for two typical stoichiometry problem types, Mass-Mass and Volume-Mass.

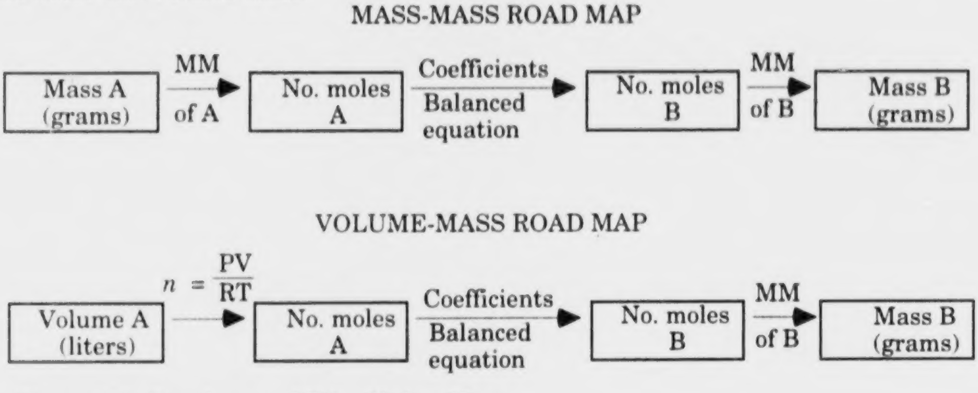

MASS-MASS ROAD MAP
Mass A (grams)
MM of A
No. moles A
Coefficients Balanced equation
No. moles B
MM of B
Mass B (grams)
VOLUME-MASS ROAD MAP
Volume A (liters)
n = PV/RT
No. moles A
Coefficients Balanced equation
No. moles B
MM of B
Mass B (grams)

Figure 30. Roadmaps. MM = Molar Mass

Another alternative to the Step 4 setup on the previous page is:

Line 1 Mass unknown, mass known
Line 2 MM unknown, MM known
Line 3 Balanced equation
Line 4 Moles unknown, moles known

This way automatically sets up the terms $(\frac{\text{mass}}{\text{MM}})$ for finding moles of known.

Some teachers (and students) prefer the result of handling the entire problem with **dimensional (or unit) analysis**. If 512 g sulfur burn completely in oxygen the product could be found as follows:

$$512\text{ g } S_8 \times \frac{1\text{ mol } S_8}{256\text{ g } S_8} \times \frac{8\text{ moles } SO_2}{1\text{ mole } S_8} \times \frac{64\text{ g } SO_2}{1\text{ mole } SO_2} = 1024\text{ g } SO_2$$

The problem worked in Part B, Steps 4-7 above could be done in steps as:

$$10\text{ mol } O_2 \times (2\text{ mol } S_8/16\text{ mol } O_2) = 20/16\text{ mol } S_8 \text{ used}$$

$$20/16\text{ mol } S_8 \times (256\text{ g/mol } S_8) = (20/16) \times 256\text{ g } S_8 \text{ used}$$

$$20/16\text{ mol } S_8 \times (8\text{ mol } SO_2/\text{mol } S_8) \times (64\text{ g/mol } SO_2) = 640\text{ g } SO_2$$

This set of steps can also be combined into a single equation as:

$$10\text{ mol } O_2 \times (2\text{ mol } S_8/16\text{ mol } O_2) \times (256\text{ g/mol } S_8 \times 64\text{ g/mol } SO_2) = 640\text{ g } SO_2$$

Students uncomfortable with the mole or unit factor-conversion approaches may prefer to use ratios. For example, in the sulfur-oxygen reaction, knowing that 1 mole SO_2 has a molar mass of 64 grams, for 10 moles SO_2, we set up the ratio:

$1\text{ mol}/ 64\text{ g} = 10\text{ mol}/x\text{ g}$ Solving for x, $x = 10 \times 64 = 640\text{ g}$

Considering the question "How much SO_2 is formed when 512 g of S_8 burn?" It is a good idea to demonstrate that trying to use the equation coefficient ratios with the mass fails to produce a correct answer.

$$\overset{512\text{ g}}{S_8 + 8O_2} \rightarrow \overset{x\text{ g}}{8SO_2}$$

$$\frac{512\text{ g } S_8}{1\text{ mol } S_8} = \frac{x\text{ g } SO_2}{8\text{ mol } SO_2}$$

$$x = 8\,(512\text{ g}) = 4096\text{ g } SO_2$$

This is in contrast to the correct value 1024 g SO_2.

The **ratio or proportion method** was popular a few decades ago and may still be found in textbooks. It suffers from not producing the value for moles of substances under consideration thus preventing quick generation of additional information about other substances in the reaction, or perhaps, information concerning volumes rather than mass. An example follows:

$$\overset{512\text{ g}}{S_8} + 8O_2 \rightarrow \overset{x\text{ g}}{8SO_2}$$

$$x = 1024\text{ g } SO_2$$

$$\frac{\text{Mass } S_8}{1\,(\text{MM } S_8)} = \frac{\text{Mass } SO_2}{8\,(\text{MM } SO_2)}$$

$$\frac{512\text{ g}}{1\text{ mol }(256\text{ g/mole})} = \frac{x\text{ g } SO_2}{8\text{ mol }(64\text{ g/mole})}$$

$$\frac{512\text{ g } S_8}{1\text{ mol } S_8} = \frac{x\text{ g } SO_2}{8\text{ mol } SO_2}$$

It is also necessary to show students problems where the mass of one reactant or product is used to find other reactants or products, respectively; otherwise, the students presume the process is only for finding products from known quantities of reactants or *vice-versa*.

Additional Practice Problem

Another good idea is to work out the quantities of all of the substances in some complex reaction and demonstrate the conservation of mass. This approach is helpful because it gives students a method of internally verifying their calculations. A sample reaction is shown below with a set of answers.

One step in the conversion of phosphate rock to fertilizer is the reaction:

$$2Ca_5F(PO_4)_3 + 7H_2SO_4 \rightarrow 2HF + 3Ca(H_2PO_4)_2 + 7CaSO_4$$

If 100 g of the phosphate compound (mineral name fluorapatite) are dissolved for laboratory analysis, how much sulfuric acid will be needed and how much of each product results (Figure 31)?

Substance	**$Ca_5F(PO_4)_3$**	**H_2SO_4**	**HF**	**$Ca(H_2PO_4)_2$**	**$CaSO_4$**
Mass (g)	100	68.0	3.96	68.9	94.4
MM	504	98.0	20.0	232	136
Moles	0.198	0.694	0.198	0.297	0.694

Figure 31. Quantities of reactants and products.

SOLUTION CONCENTRATIONS: MOLARITY

Students are familiar with dilute sugar solutions, in which very little sugar is added to a drink, in contrast to more concentrated sugar solutions. The various types of solutions, and the different ways of expressing their concentrations, are detailed in the *Content in a Nutshell, Central Concepts* sections, and Demonstration 1 of the *Solutions* module. If students are somewhat familiar with the kinetic molecular theory of gases, it may help them to see the demonstration suggested in Figure 32. In a one-liter graduated cylinder, a brown gas (NO_2) is generated by the action of a small amount of concentrated nitric acid on small pieces of thin copper wire. At the same time, a small amount of methyl orange solid is introduced into a large amount of water, in a separate one-liter graduated cylinder.

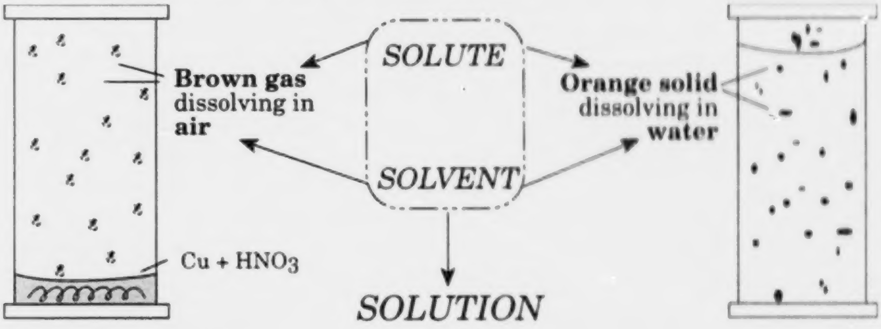


Figure 32. Solutions.

The top of the nitric acid-copper containing cylinder is covered with glass or plastic to prevent NO_2 from escaping. Although the brown NO_2 gas rises to eventually fill the graduated cylinder, the methyl orange first sinks to the bottom and then fills the graduated cylinder as it spreads out in the water. There is enough similarity between the two processes to help students picture both as solutions: in one, a gas (NO_2) dissolves in another gas (air); in the other, a solid (methyl orange) dissolves in water. Of course, air is not the pure solvent water is, and you may wish to talk about this.

Students should realize that solution concentrations can be and are expressed in various ways: grams of solute per 100 grams of solution; parts per million (ppm), really parts solute per million parts of solution. Most familiar to chemists is molar concentration, symbolized by M and defined as moles of solute per liter of solution. This discussion is limited to the common solution of a solid dissolved in water.

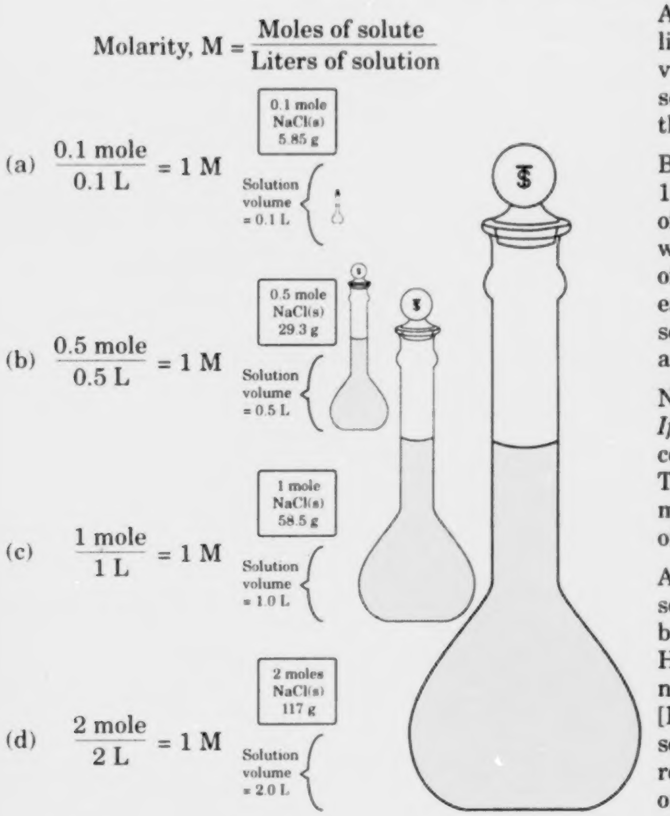


Figure 33. Molarity review.

As illustrated in Figure 33, molarity, M, is a ratio, like density, except that the units are in moles per volume rather than grams per volume. Thus *all* solutions pictured in Figure 33 are 1 M, even though they do not all contain 1 mole of dissolved solute.

By definition, each flask contains a 1M solution. In (a), 10 of those flasks would contain 10 tenths, or 1 mole of NaCl in 1 liter of solution. In (b), two of those flasks would contain two halves, or 1 mole of NaCl in 1 liter of solution. If (d) flask were divided into two halves, each half would contain 1 mole of NaCl in 1 liter of solution. If each flask contained only 1/10th (0.1) the amount of NaCl shown, each would be a 0.1 M solution.

Note that only Solution (c) has 1 mole of solute NaCl. *If* one liter of each of the other solutions were considered, they too would each contain 1 mole. Thus, both molarity and volume determine how many moles of the solute are needed. But the *ratio* of moles to volume for *any* 1 M solution *must* be 1.

As indicated in Figure 33, the number of moles of solute required to prepare any solution depends both on the required molarity and the needed volume. Here, units and unit canceling help. Suppose you need to know how many grams of lead nitrate [$Pb(NO_3)_2$] will be needed to prepare 25 mL of 0.5 M solution. The 0.5 M tells that 0.5 moles would be required to make 1.0 liter of solution. The 25 mL is only 25/1000, or 1/40 liter, so the number of moles needed for this volume is 0.5, or 1/2 mol/L x 1/40 L = 1/80 mol = 0.0125 mol. Finally, this 0.0125 mol must be converted to grams. Thus for solutions: Moles solute = M(mol/L) x V(L)

When actually preparing solutions, students need to realize that the required mass of solute must be added to *less* than the needed volume, since molarity involves volume of *solution*. Finally, distilled water of the appropriate solvent must be added to achieve this total required volume.

While molarity (M) is by far the most useful unit for solution concentration calculations, others are widely encountered. Some are described here.

Molality

Molality (m) is defined as the number of moles of solute per 1000 grams of solvent. When water is the solvent and solute concentrations are low, this is nearly equal to molarity. Calculation of colligative properties (osmotic pressure, vapor pressure, and differences in boiling and melting points from those of pure solvents) require this concentration unit. To differentiate the symbols when using script, chemists make a pointed *M* for molarity and a rounded *m* for molality so the shape will differentiate even if the size of the letter is questionable.

Percent

Percent concentration is understood to be the number of grams of solute per 100 grams of solution unless percent by volume or mole is specified. Alcoholic beverage percentages are given by volume. Wines are usually 12% by volume, that is 12 mL ethanol per 100 mL wine. Mole percent is moles solute divided by total number of moles of solvent and solutes present. All these measures are expressible as fractions (weight fraction, mole fraction, volume fraction) by simply not multiplying by 100 to create percent.

Parts per Million

Newspapers and environmental reports often use parts per million (ppm) as a concentration unit. This unit refers to the number of particles (atoms, ions, molecules) per million total particles (solutes and solvent) in the system. Similar units of parts per billion (ppb), parts per thousand (ppt), and parts per hundred (pph) are used when appropriate. One would say 10 pph rather than 100 000 ppm. Note that pph is the same as percent.

Consider the following to compare various expressions of concentration for the same solution. If 4.00 g NaOH (0.100 mol) were dissolved in 99.00 g water (5.500 mol) to make 100.00 mL solution the concentration listed in Figure 34 would result. Note that volume percent was not included because the solute was solid rather than liquid.

Concentration Unit	Calculation Expression	Concentration of Solute
Molarity	0.100 mol/0.100 L =	1.00 M
Molality	0.100 mol/0.099 kg =	1.01 m
Weight Fraction	4.00 g/103.00 g =	0.0388
Weight Percent	0.0388 x 100 =	3.88 %(wt)
Parts per thousand	0.0388 x 1000	38.8 ppt
Parts per million	0.0388 x 1 000 000	38800 ppm
Parts per billion	0.0388 x 1 000 000 000	38800000 ppb
Mole Fraction	0.100 mol/5.6 mol=	0.0178
Mole Percent	0.0178 x 100 =	1.78 %(mol)

Figure 34. Comparison of concentration values for a solution.

IONIC AND NET IONIC EQUATIONS

Most reactions students encounter in beginning chemistry courses are between solutions of ionic solids dissolved in water. Students should see demonstrations of the conductivity of NaCl(aq) *vs.* table sugar solution. The former will conduct, the latter will not. Since conductivity involves moving charged particles, the equation for salt dissolving in water,

$$\mathrm{NaCl(s)} \rightarrow \mathrm{Na^+(aq)} + \mathrm{Cl^-(aq)}$$

shows the charged particles (the ions) as well as no net charge in the solution.

An effective way to convince students that ionic solutions have two sets of properties, because they have two sets of dissolved particles, is to take equal volumes of the following 0.1 M solutions: $CuSO_4$(aq), $Cu(NO_3)_2$(aq), $CuCl_2$(aq), NaCl(aq), HCl(aq). Point out that the first three solutions are all blue because of the common Cu^{2+} ion they each contain, while the last three all form white precipitates when $AgNO_3$(aq) is added. This must be due to the common Cl^-(aq) ion in the last three. Thus $CuCl_2$ (aq) is blue due to Cu^{2+}(aq) ion and able to form a precipitate with $AgNO_3$(aq) because of the Cl^-(aq) ion.

Students should also realize that the reaction between NaCl(aq) and $AgNO_3$(aq) can be represented in several ways. The molecular equation is,

$$\mathrm{AgNO_3(aq)} + \mathrm{NaCl(aq)} \rightarrow \mathrm{AgCl(s)} + \mathrm{NaNO_3(aq)}$$

where AgCl(s) represents the white precipitate. Since both original solutions conduct, they must contain ions. Hence, a better representation is the ionic equation, where proper formulas and charges can be located in a table of common ions (Figure 36). The ionic equation,

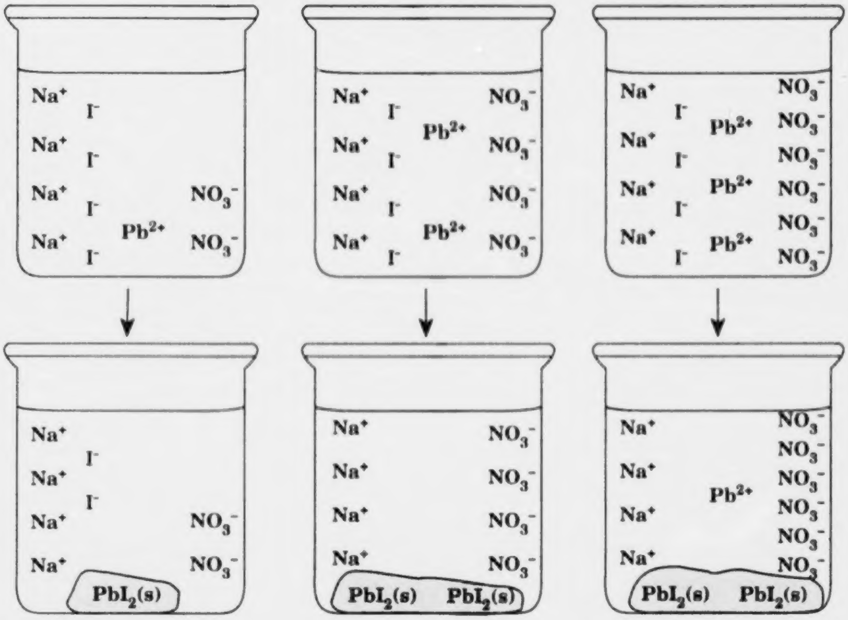
$$Ag^+(aq) + NO_3^-(aq) + Na^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)$$

shows that the $Na^+(aq)$ and $NO_3^-(aq)$ ions remain unchanged after reaction. They are therefore given the name *spectator ions*. The reaction is between the $Ag^+(aq)$ and $Cl^-(aq)$ ions to form the precipitate AgCl(s). The net change is thus best represented by what is known as the net ionic equation:

$$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$$

Why certain oppositely charged ions join together to precipitate, like the $Ag^+(aq)$ and $Cl^-(aq)$, while other oppositely charged ions remain apart, dissolved in the solution [the $Na^+(aq)$ and $NO_3^-(aq)$ for example], is an intriguing one. If enough different solutions are mixed and studied, it is easy to determine which pairs of ions always precipitate. Such determinations are summarized in solubility rules. But being able to predict when a precipitate will form, as well as its formula, is not the same as understanding *why* only certain ions do this. We must consider the energies of hydration of the separated ions contrasted with the bond strength of the solid and relative ionic sizes to answer this question. This topic is beyond the scope of most introductory courses.

The reaction between ions in lead nitrate [$Pb(NO_3)_2$] and sodium iodide (NaI) solutions, similar to the reaction in Demonstration 3 of *Simple Chemical Reactions* module, is illustrated for three cases in Figure 35. Since each $Pb^{2+}(aq)$ ion can combine with two $I^-(aq)$ ions to form the neutral $PbI_2(s)$ precipitate, excess I^- or Pb^{2+} ions simply remain unreacted (left- and right-hand boxes). Only the middle boxes or "beakers" have no excess of either, so only the spectator Na^+ and NO_3^- ions remain unreacted after reaction. As before, these ions are spectators—they never precipitate. To save space, the (aq) that should properly follow the formula and charge of each ion has been omitted. The upper diagrams in Figure 35 represent the solution before reaction; the lower diagrams represent the solution after reaction.



Excess NaI(aq) *No excess of either* *Excess $Pb(NO_3)_2(aq)$*

Molecular equation: $Pb(NO_3)_2(aq) + 2NaI(aq) \rightarrow PbI_2(s) + 2NaNO_3(aq)$

Ionic equation: $Pb^{2+} + 2NO_3^- + 2Na^+ + 2I^- \rightarrow PbI_2(s) + 2Na^+ + 2NO_3^-$

Net ionic equation: $Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

Figure 35. Reaction of lead nitrate and sodium iodide.

Since ionic compounds must be electrically neutral, students can construct proper formulas from the ionic charges. Thus Al^{3+} and OH^- must form the compound $Al(OH)_3$, while Mg^{2+} and Cl^- form $MgCl_2$. Some drill may be helpful. Once the Periodic Table has been studied, students should be able to predict charges of simple ions of representative elements. All alkali metals, for example, form stable +1 ions: Na^+, K^+, *etc.* All alkaline earths form stable ions by losing two electrons: Mg^{2+}, Ca^{2+}, *etc.* Similarly, halogens become more stable by gaining an electron: F^-, Cl^-, Br^-, I^-. This approach will not work for complex ions, such as $SO_4{}^{2-}$ and $NO_3{}^-$, or for ions of the transition metals. Figure 36 gives the names and formulas of common ions.

Cations

+1 Charge		**+2 Charge**		**+3 Charge**	
Formula	**Name**	**Formula**	**Name**	**Formula**	**Name**
$NH_4{}^+$	Ammonium	Ba^{2+}	Barium	Al^{3+}	Aluminum
Cu^+	Copper(I)	Cd^{2+}	Cadmium	As^{3+}	Arsenic(III)
H^+	Hydrogen	Ca^{2+}	Calcium	Cr^{3+}	Chromium(III)
Li^+	Lithium	Co^{2+}	Cobalt(II)	Fe^{3+}	Iron(III)
K^+	Potassium	Cu^{2+}	Copper(II)		
Ag^+	Silver	Fe^{2+}	Iron(II)		
Na^+	Sodium	Pb^{2+}	Lead(II)		
		Mg^{2+}	Magnesium		
		Mn^{2+}	Manganese(II)		
		Hg^{2+}	Mercury(II)		
		Zn^{2+}	Zinc		

Anions

–1 Charge		**–2 Charge**		**–3 Charge**	
Formula	**Name**	**Formula**	**Name**	**Formula**	**Name**
$C_2H_3O_2{}^-$	Acetate	$CO_3{}^{2-}$	Carbonate	$AsO_4{}^{3-}$	Arsenate
Br^-	Bromide	$CrO_4{}^{2-}$	Chromate	N^{3-}	Nitride
Cl^-	Chloride	$Cr_2O_7{}^{2-}$	Dichromate	$PO_4{}^{3-}$	Phosphate
CN^-	Cyanide	O^{2-}	Oxide		
F^-	Fluoride	$SO_4{}^{2-}$	Sulfate		
I^-	Iodide	S^{2-}	Sulfide		
OH^-	Hydroxide	$SO_3{}^{2-}$	Sulfite		
$NO_3{}^-$	Nitrate	$S_2O_3{}^{2-}$	Thiosulfate		
$NO_2{}^-$	Nitrite				
$MnO_4{}^-$	Permanganate				
$HCO_3{}^-$	Hydrogen carbonate (bicarbonate)				

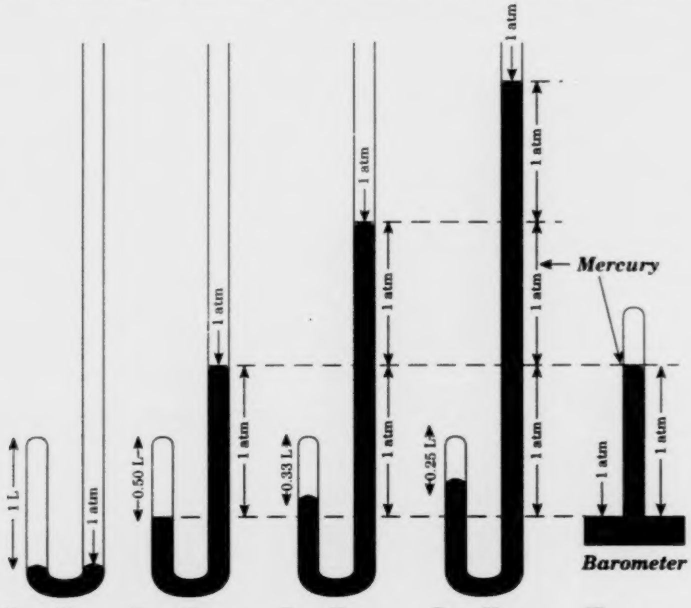
Figure 36. Names and formulas of common ions.

Mathematics Used in Chemistry

RATIO AND PROPORTION

Ratios of the form 1 is to 10 as x is to 100 should be familiar to students. These expressions are easily solved by setting up the ratios as two fractions and cross multiplying: $1/10 = x/100$, so $10x = 100$, $x = 10$. This example of a *direct* proportion leads to the definition: as one variable increases, the other increases proportionately. Obvious examples involve density, molarity, and some gas law problems. The density of water is 1 g/mL, regardless of the volume being considered. Ten mL of water has a mass of 10 g; 100 mL, 100 g; and so on. The ratio of mass divided by volume is always constant; in this case it is one. Molarity is very similar, as shown in Figure 33. A final example of direct proportion is the relationship between gas volume and Kelvin temperature (if pressure and amount of gas remain constant), or between a gas pressure and Kelvin temperature (volume and amount of gas being constant). As an example of the latter, consider the pressure of a confined gas (at constant volume) is 1.0 atm at 27 °C. Find its new pressure if temperature is raised to 327 °C (volume unchanged). Convert °C to K by adding 273. The two temperatures are 27 + 273 = 300 K and 327 + 273 = 600 K. Pressure is directly proportional to Kelvin temperature. Since temperature doubles, so must pressure. Hence the new pressure must be 2.0 atm.

Inverse proportions, where one variable decreases while the other increases, are generally more difficult for students. Boyle's Law problems are examples. Where an initial pressure, P_1 and an initial volume, V_1, of a given amount of gas at constant temperature, are involved, as shown in Figure 37, one variable times the other equals a constant. Here, $P_1V_1 = P_2V_2 = P_3V_3$, *etc.* because they all equal the *same* constant, k. The setup shown in Figure 37 might be demonstrated with a long, strong **J**-tube.



$P_1 \times V_1 =$ 1 atm x 1 L = 1 L-atm

$P_2 \times V_2 =$ 2 atm x 0.50 L = 1 L-atm

$P_3 \times V_3 =$ 3 atm x 0.33 L = 1 L-atm

$P_4 \times V_4 =$ 4 atm x 0.25 L = 1 L-atm

Thus $P \times V = K$ = 1 L-atm

Figure 37. Boyle's Law illustrated.

A typical Boyle's Law problem is: A given mass of gas at 25 °C occupies a volume of 10 L at a pressure of 1 atm. Find its volume if pressure increases to 5 atm at 25 °C. Since temperature and amount of gas are constant, we can write:

$$P_1 \times V_1 = P_2 \times V_2$$

Substituting into this expression, we get 1 atm x 10 L = 5 atm x V_2. Solving for $V_2 = 1\ \cancel{\text{atm}} \times 10\ \text{L} / 5\ \cancel{\text{atm}} = 2\ \text{L}$.

DECIMALS, FRACTIONS, PERCENT

This arithmetic recurs in chemistry, so students need to be comfortable using it. While it is good to use decimal fractions, often fractions such as 1/2 or 1/3 are more convenient than 0.5 or 0.333 when solving problems. Emphasize that formal use of proper fractions should be restricted to counting operations that have no uncertainty because only decimal fractions are capable of conveying a measured value's uncertainty. For example, it is appropriate to say 6 atoms is 1/2 dozen, but 6.0 g carbon must be reported as 0.50 moles.

Percent, meaning parts per hundred, occurs in measurement when determining uncertainties. Sometimes, experimental strategies are dependent upon percent uncertainties. This is detailed, for example, in the *Data Analysis and Concept Development* section (Teacher's Version) of the Heat of Combustion of a Candle experiment, *Thermochemistry* module. Ask your mathematics colleagues to suggest computer programs for percent-decimal fraction problems for students unsure of this area of arithmetic.

GRAPHING

Some students frequently have poor graphing skills. Their experience is often limited to preparing graphs from equations in algebra. The aspects of converting data to graphs and interpreting these graphs have seldom been encountered. An early diagnostic test of graphing ability is a wise activity. Some students need review of such basic operations as plotting a point from its coordinates.

The following guide deals with aspects of graphing related to treatment of experimental data and is probably necessary instruction for beginning students.

The procedure for treating laboratory results begins with preparing a data table. A properly designed experiment has an investigator-controlled variable called *independent* and a variable, called *dependent*, that is influenced by changes in the independent variable. The data table should have a labeled column for each variable. Usually the independent variable is listed first. Thus, each line contains a pair of values that define a data point. The labels should specify the quantity being measured and its unit, for example, time (sec) or concentration (M). The quantities should not be abbreviated but the units may be. The table should also be titled in a manner that gives more information about the experiment. The title "Determination of blood oxygen during exercise" might be appropriate for the variables suggested above.

In mathematics classes, negative numbers are common so graphs show all four quadrants. Students should be told that they will generally show only the first quadrant since they are likely to have only positive numbers in their data. There are exceptions (Celsius temperature, for example), and students will need to be reminded when negative coordinates are expected.

To create the graph, one must identify the range of values of each variable and set the axes' number lines accordingly. Students are accustomed to having a visible origin (0,0) and must be retrained to scale appropriately so data covers the full size of the graph. It must also be emphasized that each square on an axis must represent the same value. Many students will simply plot each sequential data point at the next available position without regard to the data interval. Attention must also be given to determining how rectangular graph paper is oriented to best accommodate the data, recalling that the independent variable goes on the X-axis (horizontal) and that the wider range of values requires more space for accurate plotting. The axes should be labelled and the graph titled as the table was.

If concentration were plotted *vs.* time (concentration depends on elapsed time), time would be on the X-axis. Another way of stating this is that concentration is a function of time.

After the graph layout is completed the data are plotted and a *best fit curve* drawn. This means that a smooth, possibly curving, line is drawn so that it comes as close to the points as possible and indicates the general trend of the data. It should extend only slightly beyond the extreme data points. It is unlikely that all points will fall on the line. Those that do not should be distributed uniformly around the line as in Figure 38. Examples of common student errors are shown for comparison. When the points come closer to falling on the line, they suggest that a high degree of confidence in the experiment is justified.

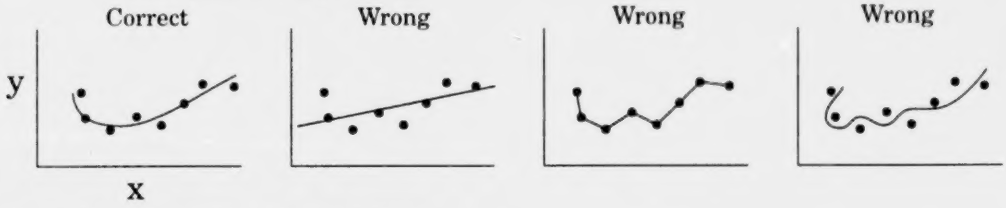
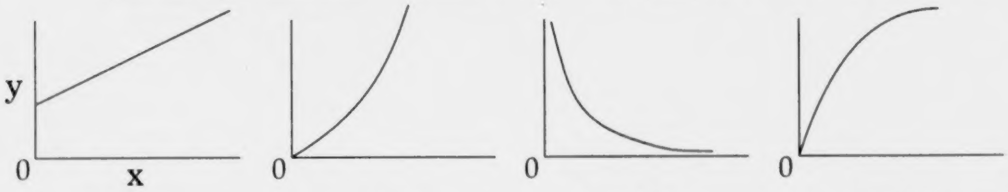


Figure 38. Examples of curve fitting.

While any shape curve might be found on a graph, the relationships studied in beginning chemistry are restricted to only a few. The simple graphs of straight line, hyperbola, parabola and logarithm with their equations are shown in Figure 39.



Straight Line
$y = mx + b$
m = slope ($\Delta y/\Delta x$)
b = y intercept

Parabola
$y = ax^2$
a = constant

Hyperbola
$xy = c$
c = constant

Logarithm
$y = \log x$

Figure 39. Common curves encountered in laboratory experiments.

Graphs show a relationship or proportionality between the variables. If the general trend is for both variables to increase in value, the relationship is *direct*. If one variable decreases while the other increases, the proportionality is *inverse*. If the points are best fit by a straight line, the relationship is *linear*; otherwise it is *nonlinear*.

Sometimes it is valuable for purposes of comparison to plot more than one line on the same graph. To do so, set the axes for the variable with the widest range. Plot the data for one line and draw its curve before starting the others. Using different colors for the lines or designating all of the points for a line by circling, boxing or some other mark is useful for differentiating them. Avoid writing equations or descriptors on the lines. Use an inset key, a window on the graph, to explain the designations for each line. An example of this type plot is given in the *Periodicity* module (ionization energy *vs.* atomic number).

If one wishes to predict a value for the graphed relationship at values beyond the data range, the curve may be extended. This is called *extrapolation*. While extrapolation may be instructive, it must be stressed that the relationship may change outside the observed range, rendering the assumption invalid. An example would be attempting to predict the speed of an accelerating car after several minutes by which time the car has already reached its top speed. It is generally safe to determine a value intermediate to data points by reading the coordinates directly from the line. This is called *interpolation*.

Students can hone their skills and learn new ones by working with computer programs for graphing (see *Computers* module). Much classroom time can be saved and better quality reports generated by having such a utility available for casual use by both teacher and students.

USE OF ALGORITHMS

The ideal gas law equation, PV = nRT, is so useful that it is the obvious example of the use of formulas in chemistry. It is mentioned in various places in the *Gases* module and is referred to in most chemistry textbooks. The first step in using any formula sensibly is to know what each symbol represents. Here, the P, V, and T are pressure, volume and Kelvin temperature, respectively. The letter n represents the amount of gas (in moles), and R is a constant known as the *molar gas constant.* To get a feel for R, rewrite the equation to solve for it: R = PV/nT. Next, consider 1.00 mole of an ideal gas (one that obeys the gas laws). At standard conditions, 0 °C or 273 K, and 1.00 atmosphere pressure, its volume is 22.4 liters. To get a numerical value for R, we simply substitute these numbers in the above expression (R = PV/nT). This gives a value:

$$R = \frac{(1.00 \text{ atm x } 22.4 \text{ L})}{(1.00 \text{ mol x } 273 \text{ K})} = \frac{0.0821 \text{ L-atm}}{\text{mol K}}$$

Of course, if the pressure and volume were expressed in other units, the numerical value of R would be different. Knowing R and any three of the terms, we can find the remaining term.

As an example, consider the *Gases* module laboratory activity. Analyzing the "Pop" in popcorn. Here, the student is asked to find the pressure (P) of the water vapor (the gas) when the corn kernel pops. From the data, students can calculate values for n, V, and T, solving for P gives P = nRT/V. The trick here is to be sure to use the R value that has the proper units. In this case, since n is moles, T in kelvin, and V in liters, the above value will work because the units work out to give P in atmospheres. Looking at just the units, P = mol x L-atm/mol K x K/L, so all units cancel except for atm.

EXPONENTS AND LOGARITHMS

Exponents have been discussed earlier. Review rules for multiplying or dividing exponents. To review, $10^3 \times 10^2 = 10 \times 10 \times 10 \times 10 \times 10 = 10^5$, so $10^3 \times 10^2 = 10^{3+2}$. Thus when multiplying numbers, one adds the exponents. But $10^3/10^2 = 10 \times 10 \times 10 / 10 \times 10 = 10^1$, so $10^3/10^2 = 10^{3-2} = 10^1$. When dividing numbers, one subtracts the exponents, as seen for an earlier calculation of the mass of a single water molecule, $18/6 \times 10^{23} = 3 \times 10^{-23}$, just as $1/10^2 = 10^{-2}$. After all, 1 is 10^0, so $10^0/10^2 = 100^{-2} = 10^{-2}$.

Logarithms (logs) are exponents, as some simple examples should remind us. The $\log_{10}100 = 2$ because $10^2 = 100$. The log is the exponent, 2, to which the base, 10, must be raised to give the number, 100. The $\log_{10}1000 = 3$ because $10^3 = 1000$. When using pH, students should be reminded of logs. Since $pH = -\log_{10} [H^+(aq)]$, knowledge of logs can help them go back and forth between hydrogen ion concentration, abbreviated $[H^+(aq)]$, and the solution's pH (See *Acids and Bases* module for a pH-concentration relationship chart.)

As an example, find the pH of a 0.01 M solution of HCl(aq). Since HCl is a strong acid (completely ionized) its solution really contains a $[H^+(aq)]$ of 0.01 M. Since 0.01 M equals 10^{-2} M, the pH must be $-(-2)$, or 2. From the other direction, the pH of pure water is 7. Find the $[H^+(aq)]$ in pure water. Since pH = 7, the negative is -7, so from the definition, $[H^+(aq)]$ must be 1×10^{-7} M.

Metric Humor

The most effective humor in our largely English system of units society is to substitute metric terms for some of our classic sayings. As an example, consider the saying "a miss is as good as a mile." To make it metric, replace the English unit mile with its metric equivalent and quip:

"A miss is as good as 1.6 km."

Some other metric maxims:

1. "First down and 9.1 m" rather than "first and ten."
2. "Every 2.54 cm a king" rather than "every inch a king."
3. "More bounce to the 28.3 g" rather than "more bounce to the ounce."
4. "155 cm, eyes of blue" rather than "five foot two eyes of blue."
5. "28.3 g of prevention is worth 0.453 kg of cure" rather than "an ounce of prevention is worth a pound of cure."
6. "Give him 2.54 cm and he'll take 1609 m" rather than "give him an inch and he'll take a mile."
7. "I'll have a 114 grammer" rather than a "I'll have a quarter-pounder."
8. What did Grandpa say to Grandma as she swatted the fly? kg. (Killer gram!)

Now think up your own. Enjoy!

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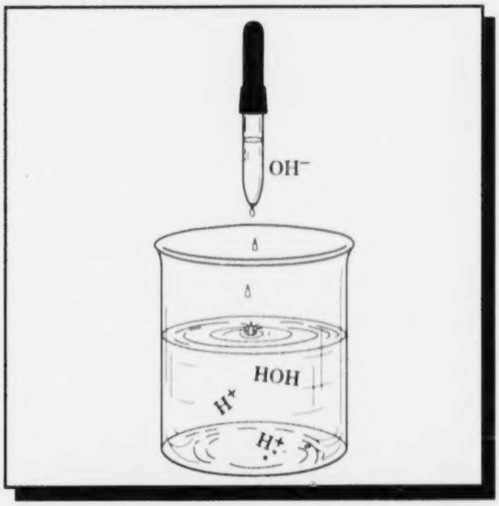


ACIDS AND BASES

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Hydrangea is a shrub with clusters of showy flowers. The color of its flowers can change as the pH of the soil in which it grows changes. The flowers are red-pink under neutral to low soil acidity and blue under conditions of higher acidity. Hydrangeas contain a natural indicator, anthocyanin, which is responsive to the pH of its surroundings. Observations such as these about hydrangeas can constitute a central theme in understanding acids and bases.

It is useful to think of an acid as a substance that releases hydrogen ions (H^+) in water solution, as proposed by Arrhenius around 1885. Since bases have always been thought of as opposites of acids, Arrhenius also proposed that bases involve hydroxide ions (OH^-). The positive hydrogen ion, which is a proton, and the negative hydroxide ion together constitute water. A more useful definition of acids and bases was proposed by Brønsted and Lowry in 1923. This definition, which gives a more complete picture of acids and bases, treated an acid as a proton donor and a base as a proton acceptor. An acid-base reaction then is essentially a transfer of protons. During the past several decades other useful definitions of acids and bases have built upon these early concepts. Some of these—the Lewis definition, for example—may be beyond the scope of the high school general chemistry classroom and less useful to high school teachers.

Neither hydrogen ions nor hydroxide ions exist alone. They are found combined with other ions of the opposite charge. Hydrochloric acid, for example, produces hydrogen ions and negative chloride ions when in solution. Sodium hydroxide, a base, produces negative hydroxide ions and positive sodium ions when in solution. When solutions of hydrochloric acid (HCl) and sodium hydroxide (NaOH) are mixed, a product resulting from the combination of sodium ions and chloride ions (NaCl) can be separated from the solution. This kind of compound is called a salt. Thus, the products of a typical acid-base reaction are a salt and water. This acid-base reaction, called neutralization, can be represented by the following equation.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + HOH(l)$$

Acids differ from one another in strength. Bases differ from one another in a similar fashion. Chemists have found that some acids produce more hydrogen ions in solution than similar amounts of others. These acids (such as hydrochloric acid, HCl, and sulfuric acid, H_2SO_4) are strong acids, whereas those that release only a few percent of hydrogen ions (hydrofluoric acid, HF, and acetic acid, $HC_2H_3O_2$) are weak acids. Similarly, sodium hydroxide, NaOH, is a strong base and aqueous ammonia, $NH_3(aq)$, is a weak base. Strong and weak acids react with strong and weak bases in neutralization reactions to produce water and salts.

It is most helpful to the chemist to be able to express quantitatively the concentration of hydrogen ion in solution. We can do this with a term called pH. At room temperature, systems below pH 7 are acidic and systems above pH 7 are basic. A sample with a pH of seven is neutral. Special indicator solutions or indicator papers impregnated with certain dyes change colors at various pH values and allow us to estimate the pH of a solution. pH meters provide a more exact measurement of hydrogen ion concentration.

Acids and bases have real-life significance. The human body functions properly only when delicate acid-base balances are maintained; crops grow best in soil with the proper pH; substances released into the atmosphere as pollutants often form acid rain. In addition, many substances used in the home are acids or bases.

PLACE IN THE CURRICULUM

Acids and bases are an important topic in high school chemistry. This module can be used qualitatively to identify substances as acids or bases, in problem solving, or as an extension for related topics such as conjugate acid-base pairs, hydrolysis, and buffers.

CENTRAL CONCEPTS

1. **Operational Definitions**. Acids and bases are chemical species that exhibit distinctive sets of observable properties. Acids taste sour (like vinegar and lemon), cause blue litmus to turn red, liberate hydrogen gas when they react with certain metals (like iron, zinc, and aluminum), and neutralize bases. Bases taste bitter (like soap), feel slippery to the touch, cause red litmus to turn blue, and neutralize acids.
2. **Conceptual Definitions**. Acids and bases can be defined conceptually to help account for what is happening on a microscopic level.

   Arrhenius concept. An acid is a substance that, when dissolved in water, forms hydrogen ions (or protons, H^+). A base is a substance that, when dissolved in water, forms hydroxide ions (OH^-). The Arrhenius concept is limited in several ways. Hydroxide ion is singled out as the only source of base character. However, other species can display basic properties (*e.g.*, ammonia). The hydrogen ion cannot exist alone in water solution. It is chemically bonded to water to form what is conventionally written as the hydronium ion, H_3O^+. In fact, the hydronium ion is associated through hydrogen bonding with a variable number of water molecules (*e.g.*, $[H_9O_4]^+$, in which H_3O^+ is associated with three water molecules). Moreover, the Arrhenius concept does not cover acid-base reactions in nonaqueous solvents.

   Brønsted-Lowry concept. An acid is a proton donor; a base is a proton acceptor. Any Arrhenius acid is also a Brønsted-Lowry acid, and any Arrhenius base is also a Brønsted-Lowry base. Hydrochloric acid (HCl) is an Arrhenius acid and is therefore, a Brønsted-Lowry acid. As a Brønsted-Lowry acid, HCl can donate a proton to water.

   $$HCl(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + Cl^-(aq) \qquad (1)$$

   Magnesium hydroxide, $Mg(OH)_2$, is an Arrhenius base and therefore also a Brønsted-Lowry base. As a Brønsted-Lowry base, OH^- can accept a proton:

   $$OH^-(aq) + H_3O^+(aq) \rightleftharpoons 2H_2O(l) \qquad (2)$$

   Ammonia, NH_3, can also serve as a Brønsted-Lowry base, because it can accept a proton from water:

   $$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq) \qquad (3)$$

   Water can function both as a Brønsted-Lowry acid (Equation 3) and a Brønsted-Lowry base (Equation 1). The Brønsted-Lowry concept of acids and bases is more general than is the Arrhenius concept. In the Brønsted-Lowry concept a base accepts protons (OH^- is only one example). Brønsted-Lowry acids and bases can be ions (*e.g.*, OH^-) or molecules (*e.g.*, NH_3 and H_2O). Some species can act as either acids or bases, depending on the nature of the other reactant (*e.g.*, H_2O). Furthermore, Brønsted-Lowry acid-base reactions are not restricted to aqueous solution. For example, in the reaction between ethyl acetate and sodium ethoxide in ethyl alcohol, ethyl acetate is a Brønsted-Lowry acid and ethoxide ion is a Brønsted-Lowry base.

   $$\underset{\text{BL acid}}{CH_3COOC_2H_5} + \underset{\text{BL base}}{Na^+O^-C_2H_5} \rightarrow {}^-CH_2COOC_2H_5 + HOC_2H_5 + Na^+$$

3. When a suitable amount of base is added to an acid solution, the base and acid properties disappear and the acid is said to be neutralized. A neutralization reaction is the reaction of an acid and a base that results in an ionic compound and possibly water. The ionic compound is called a salt.

4. Salts are made up of a cation (other than H^+) and an anion (other than OH^- or oxide, O^{2-}). The formula of the salt indicates the acid and base that could react to prepare the salt. The cation (positive ion) is derived from the base; the anion (negative ion) is derived from the acid.

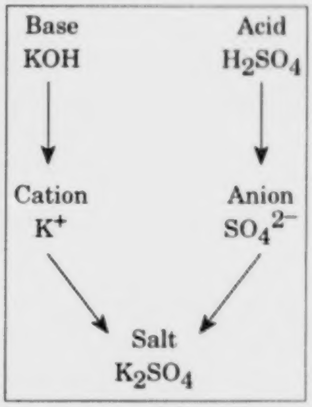


Figure 1. Formation of salt from acid and base.

5. A strong electrolyte is an electrolyte that exists in solution almost entirely as ions. Hydrochloric acid, HCl, is a strong electrolyte. A weak electrolyte is an electrolyte that dissolves in water to form an equilibrium between a molecular substance and a relatively small quantity of ions. Acetic acid and aqueous ammonia are examples of weak electrolytes.

6. In aqueous solution, H^+ (that is, H_3O^+) is the strongest stable acid, and OH^- is the strongest stable base.

7. pH is the negative logarithm of the molar hydrogen ion concentration; pOH is the negative logarithm of the molar hydroxide ion concentration. Mathematically,

$$\mathrm{pH} = -\log[\mathrm{H^+}] \quad \text{and} \quad \mathrm{pOH} = -\log[\mathrm{OH^-}]$$

8. $\mathrm{pH} + \mathrm{pOH} = 14.0$ (at 25 °C)

9. The relative acidity (or basicity) of a solution can be experimentally determined with an acid-base indicator. Indicators are either weak acids or weak bases that change colors over a short pH range. That point in a titration when a suitably chosen indicator changes color is called the *end point*. Some typical acid-base indicators, their acid and base colors, and the pH range for the color change are given in Figure 2.

Indicator	Acid color	Base color	pH range
Bromophenol blue	Yellow	Blue-violet	3.0 - 4.5
Methyl orange	Red	Yellow	3.1 - 4.4
Universal indicator	Red	Blue-violet	4.0 - 10.0
Litmus	Red	Blue	4.7 - 8.2
Methyl red	Red	Yellow	4.8 - 6.2
Bromthymol blue	Yellow	Blue	6.0 - 7.6
Phenol red	Yellow	Red	6.8 - 8.4
Phenolphthalein	Colorless	Red	8.2 - 10.0

Figure 2. Indicator colors and pH range.

10. A pH meter can give quantitative pH values to several decimal places. Usually, a specially designed combination electrode is dipped into the solution to be tested. An electrical potential is generated between the electrodes that depends on the pH. The electrical potential, in volts, is digitally displayed or read on a meter calibrated directly in pH.

RELATED CONCEPTS

1. **Ions**
   a. An **ion** is an electrically charged atom or group of covalently bonded atoms—for example, sodium ion, Na^+, and carbonate ion, CO_3^{2-}.
   b. An **electrolyte** is a substance that, in water solution, is capable of conducting an electric current. An aqueous solution of sodium chloride conducts an electric current; therefore NaCl is an electrolyte. A **nonelectrolyte** is a substance that does not dissociate into ions in solution and therefore does not conduct an electric current. Sugar is a nonelectrolyte.
   c. **Ionization of covalent substances** is the breaking apart of a covalent substance into ions in aqueous solution. Hydrogen chloride ionizes completely in aqueous solution into hydrogen ions and chloride ions and is thus classified as a strong acid. Acetic acid only slightly ionizes in water; it is a weak acid.
   d. **Dissociation** is separation of an ionic substance into ions in an aqueous solution. For example, sodium chloride, NaCl, dissociates as follows:

$$NaCl(s) \cancel{+} H_2O(l) \rightarrow Na^+(aq) + Cl^-(aq)$$

2. **Solution concentration**. The ratio of the quantity of solute to the quantity of solution (or solvent). It may be expressed as molal concentration, molar concentration, weight percent, volume percent, mole fraction, *etc.* The molarity or molar concentration (M) of a solution is the number of moles of solute per liter of solution. A solution of 20.0 g sodium hydroxide, NaOH, in one liter of solution has a concentration of 0.500 mol/L NaOH, or 0.500 M NaOH.
3. Classification and identity of some common acids and bases.

   Strong acids. There are six common strong acids; molarities of these acids as they are commonly sold are indicated in Figure 3.

Formula	Name	Molarity	Weight %	Density (g/mL)
HI	Hydriodic acid	7.6	57.0	1.70
HBr	Hydrobromic acid	9.0	50.0	1.52
HCl	Hydrochloric acid (muriatic acid)	12.0	37.0	1.18
HNO_3	Nitric acid (aqua fortis)	15.0	70.0	1.35
$HClO_4$	Perchloric acid	11.6	70.0	1.67
H_2SO_4	Sulfuric acid (oil of vitriol)	18.2	97.0	1.84

Figure 3. Common strong acids.

Weak acids. Unless otherwise in-formed, one can assume other common acids are weak. Figure 4 presents three examples:

Formula	Name	Molarity	Weight %	Density (g/mL)
$HC_2H_3O_2$ or CH_3COOH	Acetic acid (glacial acetic acid or vinegar)	17.4	99.7	1.05
HF	Hydrofluoric acid (acid of fluor)	31.3	53.0	1.18
H_3PO_4	Phosphoric acid	14.7	85.0	1.69

Figure 4. Some weak acids.

Strong bases. The hydroxides of alkali and alkaline earth elements (except Be). Two common strong bases are shown in Figure 5.

Formula	Name
KOH	Potassium hydroxide (caustic potash)
NaOH	Sodium hydroxide (caustic soda or lye)

Figure 5. Some strong bases.

Weak bases. Unless otherwise informed, one can assume other common bases are weak. One example is shown in Figure 6.

Formula	Name	Molarity	Weight %	Density (g/mL)
NH_3	Ammonia (spirit of hartshorn)	15.7	30.0	0.890

Figure 6. A weak base.

4. Equation writing and balancing
5. Stoichiometry
6. Exponents, powers of ten, logarithms

RELATED SKILLS

1. Reading volumetric devices
2. Using calculator for exponents, logarithms, arithmetic operations
3. Proper handling of potentially hazardous liquids and solids. Minimum amounts of acids and bases should be kept in the laboratory, preferably only those amounts needed for a given activity.

PERFORMANCE OBJECTIVES

After completing their study of acids and bases, students should be able to:

1. define and classify acids and bases operationally and conceptually.
2. write neutralization equations, given the identities of reacting acids and bases, and complete relevant calculations based on these equations.
3. identify the salt product produced in an acid-base reaction. Given a salt, specify the acid and base from which it could be produced.
4. complete an acid-base titration (small-scale or standard).
5. determine whether a solution is acidic or basic using an indicator or a pH meter.
6. interpret pH values in terms of powers of 10.
7. distinguish between a strong acid (or base) and a weak acid (or base) operationally and conceptually.
8. present useful mental pictures of (a) ionization, (b) dilute *vs.* concentrated acid/base, (c) weak *vs.* strong acid/base, and (d) neutralization.
9. use a matrix to record and interpret experimental data.
10. provide a chemical explanation for acid rain.
11. express their opinions about the roles of acids and bases in the world.
12. explain how their study has changed their opinions about acids and bases and other substances.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Classifying Substances Based on Their Reactions

Introduction

Chemists find it convenient to classify the overwhelming number of known substances into categories that share common properties. In this laboratory activity you will become familiar with some chemical behaviors of certain types of substances that allow them to be classified into acids, bases, and salts.

Purpose

To develop a scheme for classifying different substances as acids, bases, or salts.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. The acids and bases used are corrosive. Follow your teacher's instructions in handling acids and bases and for waste disposal. All spills should be cleaned up immediately.
3 Accidental spills on the body should be flushed with water; notify your teacher.
4. Barium hydroxide is toxic.
5. Dispose of the chemicals as your teacher directs.

Part I. Procedure

1. Arrange a matrix within the well plate or among the test-tubes to match the pattern shown in Figure 7. Add a drop or two of each test solution.
2. Add a drop of phenolphthalein to a sample of each of the seven test solutions. Record your observations.
3. Add a drop of bromthymol blue to a different sample of each of the seven test solutions. Record your observations.
4. Add a drop of universal indicator to a different sample of each of the seven test solutions. Record your observations.
5. Clean a piece of magnesium ribbon with fine-grade sandpaper or steel wool until it is shiny. Cut the ribbon into seven pieces small enough to fit in the test-tubes or wells.
6. Add a piece of magnesium, Mg, to each of the seven test solutions. Record your observations.
7. Add a drop of calcium nitrate solution, $Ca(NO_3)_2(aq)$, to a different sample of each of seven test solutions. Record your observations.
8. Use the conductivity tester on the last sample of each of the seven test solutions. Rinse the conductivity tester in fresh distilled water after testing each sample. Record your observations.

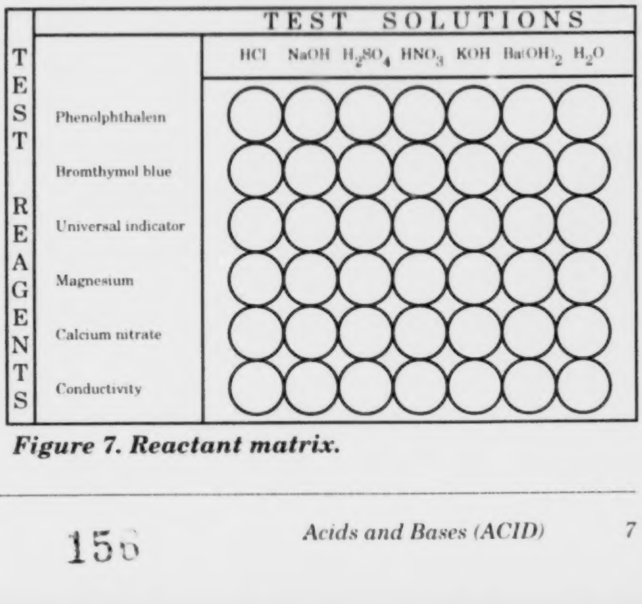


Figure 7. Reactant matrix.

Part I. Data Analysis

1. Use your observations to group the test solutions into as many categories as are needed to account for similarities and differences in chemical reactivity. Be prepared to justify your classification scheme.
2. For those solutions that indicated high conductivity with the tester, identify the ions present in solution.
3. Hydrochloric acid is a member of a class of compounds called acids. Sodium hydroxide is a member of a class of compounds called bases. What other substances in the group of seven tested should be called acids or bases?

Part I. Implications and Applications

1. What strategy could you use to identify an unknown substance as an acid or a base?
2. A drop of bromthymol blue placed in an unknown solution turns yellow. Is the solution acidic or basic?
3. Why do you think universal indicator is called an indicator?
4. What ionic species is common among the group of compounds you categorized as acids? As bases?

Part II. Procedure

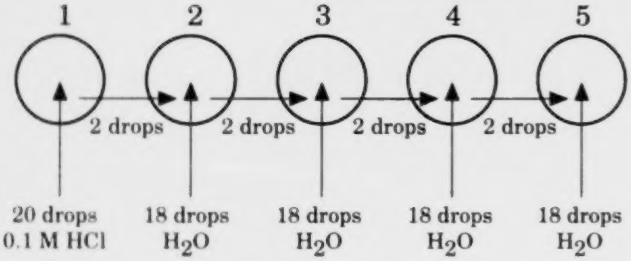


Figure 8. Serial dilution procedure.

1. Add 20 drops 0.1 M HCl to the first well or tube.
2. Remove two drops and place them in the second tube or well. Add 18 drops of water to the second tube or well, and mix.
3. Take two drops of the second solution and add them to the third well. Add 18 drops of water to the third well and mix.
4. Continue this "serial dilution" until you have five increasingly dilute solutions of hydrochloric acid, HCl, as shown in Figure 8.
5. Repeat Steps 1 through 4, starting with 0.1 M sodium hydroxide, NaOH, rather than hydrochloric acid.

		Concentration	Observations
Most Concentrated	1	0.1 M	
	2		
HCl Solutions	3		
	4		
	5		
	5		
	4		
NaOH Solutions	3		
	2		
Most Concentrated	1	0.1 M	

6. Add one drop of universal indicator to each well or tube. Record your observations.

Figure 9. Sample data table.

Part II. Data Analysis

1. Calculate the molar concentration (mol/L) of each dilution of hydrochloric acid, HCl, and sodium hydroxide, NaOH. Record these values in the chart above. If your teacher so instructs, show how you completed your calculation for the concentration of hydrochloric acid in the second well or test-tube.
2. Propose a convenient pattern to summarize your observations. (*HINT: Think about the colors of the spectrum.*)

Part III. Procedure

1. Add 3-4 drops of each commercial liquid sample to a separate well or test-tube. Label each location or draw a map so you can identify each sample.
2. For commercial solid samples, add an amount about the size of a pencil lead tip into a separate well or small test-tube. Add 3-4 drops of water to each.
3. Add a drop of universal indicator to each well or test-tube. Record your observations.
4. If your teacher so instructs, try any of the other tests you completed in *Part I* with additional samples of the commercial materials.
5. Thoroughly wash your hands before leaving the laboratory.

Part III. Data Analysis

1. Classify each commercial sample as acidic, basic, or neutral.
2. Match the particular concentration of hydrochloric acid or sodium hydroxide from *Part II* that each commercial sample most closely resembles in its acidic or basic properties.

Part III. Implications and Applications

1. How important is it for consumers to understand some fundamental ideas about acids and bases?
2. Is the acidity or basicity of a commercial product usually related to how it is used? Explain.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Classifying Substances Based on Their Reactions

Major Chemical Concepts

This laboratory activity is designed to develop operational definitions for acids, bases, and pH. It should be used as an introduction to acids and bases before students learn verbal or mathematical definitions of important terms. Conceptual definitions can be developed, if desired, during the post-laboratory discussion.

Expected Student Background

Students should be able to calculate the new concentrations of solutions after they are diluted and be familiar with standard laboratory procedures. Students should also be familiar with the electrical conductivity of ionic solutions.

Level

Appropriate for any basic or general student.

Time

This laboratory activity can be completed in two class periods. However, it can occupy over three periods if you use extended discussion to isolate, define, and solidify concept acquisition.

Safety

Read the *Safety Considerations* section in the *Student Version*. In addition, be sure that students treat commercial products as potentially hazardous; some are caustic or can lead to dangerous reaction products if mixed together. For example, if Vanish™ and household bleach are mixed together, poisonous chlorine gas is produced.

Materials (For 24 students working in pairs)

Part I. Materials

0.1 M Hydrochloric acid, HCl (2 mL conc HCl, diluted to 250 mL)
0.1 M Sodium hydroxide, NaOH (1 g NaOH, diluted to 250 mL)
0.1 M Sulfuric acid, H_2SO_4 (see *Advance Preparation*)
0.1 M Nitric acid, HNO_3 (1.5 mL conc HNO_3, diluted to 250 mL)
0.1 M Potassium hydroxide, KOH (1.5 g KOH, diluted to 250 mL)
Saturated barium hydroxide solution, $Ba(OH)_2$ [15 g $Ba(OH)_2$, diluted to 250 mL]
1.0 M Calcium nitrate, $Ca(NO_3)_2$ [8 g $Ca(NO_3)_2$ diluted to 50 mL]
Distilled water
Magnesium ribbon, Mg, 10-cm length
Bromthymol blue indicator (see *Advance Preparation*)
Phenolphthalein indicator (see *Advance Preparation*)
Universal indicator
12 Conductivity testers
12 Well-plates (96 wells) or 504 small test-tubes and racks
12 Small 25-cm^2 pieces of fine-grade sandpaper or steel wool
12 Scissors

Part II. Materials

12 Small dropper bottles of each of the following:

- 0.1 M Hydrochloric acid, HCl
- 0.1 M Sodium hydroxide, NaOH
- Distilled (or deionized) water
- Universal indicator
- 12 Well-plates or 120 small test-tubes (can clean ones used previously)
- 60 Toothpicks or small stirring rods

Part III. Materials

12 Small dropper bottles of each of the following:

- A number of commercial samples, *e.g.*, household ammonia (see *Advance Preparation*)
- Universal indicator
- Distilled (or deionized) water
- Well-plate, or small test-tube for each commercial product tested
- Conductivity tester (Directions for building a conductivity apparatus are given in the *Instrumentation* section of this module.)
- Scissors

Advance Preparation

Each student pair should have one dropper bottle of each of these:

- **0.1 M H_2SO_4**. Add 1.5 mL of concentrated (18 M) sulfuric acid, H_2SO_4, to enough water to make 250 mL of solution. *WARNING: Be sure to add acid slowly with continual stirring to several hundred milliliters of water, then make up to 250 mL of solution.*

- **Indicators**. Methyl red, bromthymol blue, phenolphthalein, and universal indicator may be purchased as previously prepared solutions from appropriate commercial vendors. Alternatively, prepare methyl red by dissolving 0.02 g in 60 mL of ethanol and 40 mL water; bromthymol blue by dissolving 0.1 g in 16 mL 0.01 M NaOH and adding 234 mL water; phenolphthalein by dissolving 0.05 g in 50/50 ethanol/water. Feenamint™ and Ex-Lax™ are alternative sources of phenolphthalein. Each group should have one dropper bottle of each indicator.

- **A group of household or commercial samples spanning a wide range of pH**. Suggestions, with approximate pH values in parentheses, include: liquid Vanish™ (pH 1), solid Vanish™ (pH 2), vinegar (pH 3), Sprite™ (pH 3), orange juice (pH 4), most vinyl cleaners (pH 4), Fluoriguard™ anti-cavity rinse (pH 5), Cepastat™ Mouth Spray (pH 5), most nail polish removers (pH 6), many shampoos (pH 6), many liquid detergents (pH 7 and above), borax solution (pH 9), window cleaner such as Windex™ (pH 10), Dow™ bathroom cleanser (pH 11), many dishwashing detergents (pH 12), Drano™ (pH 14).

Preparing a large number of solutions and filling numerous small dropper bottles with solutions is often a tedious and time-consuming task. Begin this process several days in advance of the activity. If you or your students have not already made conductivity testers (*Instrumentation* section) you should do so in advance of this activity. Be sure the electrodes are small enough to fit into the tubes or wells.

Pre-Laboratory Discussion

Little pre-laboratory discussion is needed with this activity other than a safety review and ensuring that students understand the instructions, including use of the conductivity tester. Make clear to students that they will be classifying materials based on their reactions and insure that they understand in general what "classifying" involves. *Do not* use terminology such as *acid, base, pH, etc.*; the purpose of this activity is to encourage students to invent these concepts from relevant chemical experiences.

Teacher-Student Interaction

The effectiveness of this activity will depend on your interaction with small groups as they work within parts of the procedure and with the entire class as they prepare to move from *Part I* to *Part II*, and from *Part II* to *Part III*.

Part I

Move from group to group to ensure that each group sets up the reaction matrix as instructed. If you use test-tubes rather than well plates, students may need to complete one reaction series and wash the tubes before they conduct the next reaction series. In either case, be sure students record their observations.

To prepare students for *Part II*, point out that most solutions were 0.1 M concentration. Invite students to speculate on the results if similar tests were conducted with solutions that were (a) more concentrated or (b) less concentrated. Consider recording the predictions of several students on the chalkboard before proceeding to *Part II*.

Part II

Circulate to ensure that students correctly complete the dilution series and record their observations.

When students have completed their observations, check on whether they need help in the dilution calculations.

When students have answered *Data Analysis*, Question 2, within their group, have them share their suggestions. When the gradation of colors associated with more and less acidity and less and more basicity is recognized by students, inform them that this is one way chemists identify the property called pH. Tell them that they will develop their understanding of pH as their acid-base study progresses. Universal indicator will continue to serve as a useful way to identify the pH of materials.

Point out that the idea of using universal indicator to determine pH will be used in the *Part III* of the activity. If students are to complete additional tests (in addition to universal indicator) provide suitable instructions.

Part III

As you initially circulate among students, ensure that they use only small samples of household materials and that they do not mix the materials with each other. Encourage groups to share household materials if only limited samples are on hand.

Have students share and compare information. Encourage them to use language such as "has the same color as 0.001 M NaOH" as well as "is an acid" or "is a base."

Both items in *Part III. Implications and Applications* are open-ended; encourage students to evaluate their colleagues' answers. In dealing with the second question, encourage students to relate their hypotheses to their original observations and correlate answers with the pH chart found in *Tips for the Teacher* (Item 8).

Anticipated Student Results

All acids should cause each indicator to exhibit a characteristic color; the bases should cause a different color to appear. Because water can act as both a weak acid and a weak base, its indicator's color will be intermediate to the colors shown by more concentrated and stronger acids and bases.

Magnesium metal will react with acids to produce gas bubbles, but should not react with bases. All acids and bases should be electrically conductive; water is not. Bases used in this activity should produce a precipitate (cloudiness) with calcium nitrate; acids should not.

If you are using universal indicator as recommended, the tested solutions should grade from red through yellow and green to blue and violet, following the visible light spectrum from most acidic (red) to most basic (violet).

Answers to Data Analysis

Part I

1. When analyzing data, student groups should compare classification schemes and then discuss any differences noted. After discussion, acids should be grouped together due to their common indicator colors, reactivity with magnesium, and failure to react with calcium nitrate. Bases should be grouped together based on indicator color, failure to react with magnesium, and reaction with calcium nitrate. Water should be identified as belonging to neither group. At this point, *and not before this point*, introduce the terms "Acid," "Base," and "Neutral." These terms, as used by chemists, refer to the properties (and others not yet encountered) just observed by the students.
2. The ions in solution are: $HCl(aq)$ contains $H^+(aq)$ and $Cl^-(aq)$, $NaOH(aq)$ contains $Na^+(aq)$ and $OH^-(aq)$, $H_2SO_4(aq)$ contains $H^+(aq)$ and $HSO_4^-(aq)$ [remember $HSO_4^-(aq)$ is a weak acid] and some $SO_4^{2-}(aq)$, $HNO_3(aq)$ contains $H^+(aq)$ and $NO_3^-(aq)$, $KOH(aq)$ contains $K^+(aq)$ and $OH^-(aq)$, and $Ba(OH)_2(aq)$ contains $Ba^{2+}(aq)$ and $OH^-(aq)$.
3. $HCl(aq)$, $H_2SO_4(aq)$, and $HNO_3(aq)$ are acids; $NaOH(aq)$, $KOH(aq)$, and $Ba(OH)_2(aq)$ are bases.

Part II

1. For the HCl solutions: Solution 1 is 0.1 M, 2 is 0.01 M, 3 is 0.001 M, 4 is 0.0001 M, 5 is 0.00001 M. For the NaOH solutions: Solution 5 is 0.00001 M, 4 is 0.0001 M, 3 is 0.001 M, 2 is 0.01 M, 1 is 0.1 M.
2. Follow the spectrum from red to violet as the solutions become less acidic and more basic.

Part III

1. The results depend on which household materials are tested. For general results see the pH chart in *Tips for the Teacher* (Item 8).

Post-Laboratory Discussion

The suggested class development outlined above is intended to introduce the concepts of acid, base, pH, and possibly indicators from phenomena students directly observe in the laboratory. Thus the post-laboratory discussion can employ a form of "stand-and-deliver" strategy in which you call on students to define main concepts and illustrate definitions with examples from reactions just studied. Other students can be invited to evaluate or extend previous comments by their colleagues.

Students should also be asked occasionally to give examples of acid-base phenomena that were not observed in the laboratory activity to see whether they can extend the concepts.

Answers to Implications and Applications

Part I

1. Use one of the three indicators to test for an acid or a base.
2. Bromthymol blue changes to yellow in an acid solution or in a solution of pH less than 6 (see Figure 2).
3. In solutions of different concentrations of acids or bases, the universal indicator has different colors.
4. All acids seem to have $H^+(aq)$, and all bases seem to have $OH^-(aq)$.

Part III

1. Many household products display acidic or basic properties.
2. Most commercial cleansing products are basic. Most fruits tend to be acidic.

Possible Extension

A good group activity is to conduct an acid-base scavenger hunt. Send students out with a well plate or test-tubes and universal indicator to identify five additional acids and five additional bases in their everyday environment. They can complete this as a take-home activity, if you choose.

Assessing Laboratory Learning

1. **Laboratory practical**. Give individuals or groups of students one or more unknowns. Have students categorize the samples as acidic and/or basic, indicating which solution(s) from *Part II* had comparable pH values.
2. **Demonstration examination**. Same as laboratory practical, but you demonstrate the tests. Carry out a series of reactions, describing at each step what you are adding and allowing students to observe the results. Then have students categorize and/or identify the pH of the unknown substance.
3. **Written examination**. Here are some possible items:
    a. Three identical samples of a clear, colorless liquid are tested: Magnesium is added to Sample 1; no reaction is observed. To Sample 2 is added calcium nitrate; the liquid turns cloudy. Universal indicator is added to Sample 3; the indicator turns bluish-purple. Is the unknown liquid acidic, basic, or neutral? *[Basic]*

    b. Universal indicator can be made to turn various colors, depending with what it is mixed. What property determines the color shown by universal indicator? *[Whether the substance being tested is an acid or a base.]*

    c. In what ways are some materials commonly available in grocery stores similar to materials commonly used in a chemistry laboratory? *[Some materials are acids (vinegar, 7-Up, orange juice); some are bases (baking soda, milk of magnesia).]*

Suggestions for Other Laboratory Activities

These suggested activities are in outline format. Details for starred (*) activities are found in Shakhashiri, *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (see *References*). Suggested demonstrations in the section that follows could also be used as laboratory activities. *CAUTION: Use appropriate safety guidelines in performing these activities, since only brief procedural steps are described here.*

1. **Properties of acids and bases***. Arrange eight beakers in two sets of four. Add 20 mL each of the following to separate beakers: water, 2 M HCl, 2 M H_2SO_4, and 2 M $HC_2H_3O_2$. Test each solution with litmus paper. Add several magnesium turnings, Mg, to each beaker in one set of the four liquid samples. To each beaker in the second set of four, add 10 drops of litmus solution. Pour 25 mL 6 M NaOH in each beaker. Add magnesium turnings to each beaker and compare the results with the previous set.
2. **Define acid, base, and pH operationally**. This objective can be accomplished with cabbage juice and/or universal indicator in household materials, hydrochloric acid, and sodium hydroxide at various concentrations, as outlined: Decant the purple liquid from canned red cabbage. Add 5 mL of the extracted juice to (1) 125 mL vinegar, (2) 125 mL household ammonia, (3) 1 teaspoon baking soda in 125 mL distilled water, (4) 125 mL colorless, carbonated beverage, (5) 5 mL of laundry detergent in 125 mL distilled water, (6) 125 mL milk, (7) 125 mL 0.1 M HCl, and (8) 125 mL 0.1 M NaOH. Record the color of each mixture. Decide on the indicator color of red cabbage juice in acidic and basic solutions. Repeat with universal indicator.
3. **Neutralization**. Titrate 20 mL 0.1 M HCl with 0.1 M NaOH using phenolphthalein as indicator. Repeat with 0.1 M H_2SO_4. Determine which titration requires more base. Explain results using balanced equation for neutralization of each acid by NaOH.
4. **Titration of an antacid***. Weigh an antacid tablet accurately to nearest 0.01 g. Place tablet in an Erlenmeyer flask and add 100 mL of 0.15 M HCl to dissolve tablet. Add 5 drops of bromophenol blue indicator solution. If the indicator is blue, add HCl until it turns yellow. Record volume of HCl added. Titrate the sample with 0.1 M NaOH until the solution just turns blue. Calculate the mass of HCl solution consumed per gram of antacid tablet.
5. **Distinction between acid strength and acid concentration***. Separate beakers are filled with 40-mL samples of 0.1 M HCl, H_2SO_4, and $HC_2H_3O_2$. Ten drops of phenolphthalein are added to each. The pH of each solution is determined with a pH meter or with pH test paper. 0.1 M NaOH is slowly added to each beaker of acid until the indicator just turns from colorless to pink. Compare the volumes of NaOH needed for each acid. Explain the results.
6. **Strong vs. Weak Acids***. Add 25 mL of 6 M, 1 M, 0.1 M HCl (strong acid); 6 M, 1 M, 0.1 M $HC_2H_3O_2$ (weak acid) to separate beakers. Measure pH. Check conductivity of solutions and rank.Add Mg turnings to each beaker. Rank by reactivity.
7. **Acid and Base Indicators***. Mix 10 mL indicator solution with 50 mL colorless solutions having pH values ranging from 1 to 13. For indicator solutions that change color below pH 7 use bromophenol blue or methyl orange. For indicator solutions that change color near pH 7 use litmus or phenol red. For indicator solutions that change color above pH 7 use phenolphthalein. Compare colors in solutions of various pH.

8. **pH of Familiar Products***. Prepare standards for comparison by adding 10 drops of universal indicator to 50 mL of standard solutions representing pH values of 1, 3, 5, 7, 9, and 11. Add 10 drops of universal indicator to 50 mL of familiar products (fruit juice, vinegar, carbonated colorless beverage, milk, mouthwash, ammonia, bleach). Dissolve 5 mL (one teaspoon) of any of the following in 100 mL water: laundry detergent, hand soap, shampoo, drain cleaner, baking soda, antacid, aspirin. Add 10 drops of universal indicator to 25 mL of each. Compare colors for pH. Use a pH meter to determine accurately.

9. **Concept of pH**. Serial dilution of an acid; measure pH with (1) pH paper, (2) pH meter, (3) various indicators, and/or (4) with assistance of a computer interface.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Metal-Acid Reactions

This demonstration illustrates that most active metals replace hydrogen from an acid.

Materials

6 M Hydrochloric acid, HCl
Small samples of any four metals (Cu, Mg, Al, Zn, Fe, or Sn)
Slant-top overhead apparatus or Petri dishes (or 50-mL beakers or titer plates)

Safety

6 M HCl is corrosive and must be handled with care. Goggles should be worn.

Directions

Place vessels on an overhead projector. Add 6 M HCl to each vessel, followed by the sequential addition of the metal samples. Observe; ask students to explain the results.

Demonstration 2: Nature of Indicators

Many common foods have natural indicators.

Materials

Natural indicators: Red cabbage juice, grape juice, or fillings from berry fruit pies
Vinegar
Baking soda (sodium bicarbonate), $NaHCO_3$
6 Petri dishes, slant test-tube apparatus, or 50-mL beakers

Safety

Most of the materials are very weak acids or bases, but goggles should still be worn.

Directions

Place natural indicators in each of three dishes, test-tubes, or beakers. You might want to dilute the fruit fillings with a little distilled water and use the resulting mixture as the sample. Add vinegar to one solution, add baking soda to another, and use the third as a control. Have students make observations and state their conclusions.

Demonstration 3: Indicators and the pH Scale

Different indicators change colors at different pH values.

Materials

Universal indicator solution
Bromthymol blue indicator solution
Phenol red indicator solution
Other indicator solutions if desired, or solutions of known pH: 1, 4, 7, 10

Safety

Goggles should be worn even though these solutions are weak acids and bases.

Directions

Line up twelve Petri dishes (or small beakers) on the overhead projector as shown in Figure 10.

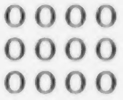


Figure 10. Arrangement of Petri dishes.

Place a separate solution of pH 1, 4, 7, and 10 in each row. Add universal indicator to each Petri dish in the first column, bromthymol blue to each dish in the second column, and phenol red to each dish in the third column. Add one to three drops (or the amount of indicator needed to produce an observable color) to each Petri dish. Either invite students to make their own observations and interpretations or, alternatively, interpret the results for them.

Demonstration 4: Conductivity

Acids and bases conduct an electric current.

Materials

Conductivity apparatus (see *Instrumentation*)
Beaker, 250-mL
10 Beakers, 50-mL
0.01 M or 0.1 M Sulfuric acid, H_2SO_4
0.01 M or 0.1 M Barium hydroxide, $Ba(OH)_2$
Phenolphthalein solution, 1%
0.1 M Hydrochloric acid, HCl
0.1 M Acetic acid, $HC_2H_3O_2$
0.1 M Aqueous ammonia, NH_3
0.1 M Sodium hydroxide, NaOH
Glacial acetic acid, $HC_2H_3O_2$
Selected household products (*e.g.*, Drano™, vinegar, *etc.*)

Safety

Many commercial products can be corrosive to the skin or damage clothing. Apron and goggles should be worn.

Conductivity Test Directions

Test the 0.1 M solutions and selected household products for their relative conductivities (do not test H_2SO_4 or $Ba(OH)_2$; save these for the conductivity titration). Use a clean beaker for each solution. Test glacial acetic acid last in this manner: test its conductivity, then double the volume by adding distilled water and test the conductivity again. Continue in this fashion until adding water no longer changes the conductivity of the acetic acid.

Conductivity Titration Directions

For chemical and electrical reasons this activity should be done only as a teacher demonstration. Set up clean burets using a ring stand and buret clamp. Fill one buret with H_2SO_4 solution and the other with $Ba(OH)_2$ solution. It is not necessary to fill each buret completely, but at least 40 mL of each solution should be present. Add 30-40 mL of H_2SO_4 from the buret to the 250-mL beaker.

Add two to three drops of phenolphthalein indicator solution. Stir. (If available, a magnetic stirrer will be helpful.) Test the conductivity. (It is best to use a conductivity tester with a 60-W light bulb for this and to leave the electrodes immersed in the liquid.) Slowly add $Ba(OH)_2$ solution from the other buret with stirring. Note the conductivity as the base solution is added. Stop the $Ba(OH)_2$ addition as close as possible to the neutrality point. The light bulb should glow more dimly until it completely stops glowing. This is the neutrality point. The phenolphthalein color should become visible at the same point.

After students have time to observe and record their observations, add additional $Ba(OH)_2$ solution. The light bulb should begin to glow increasingly brightly and the phenolphthalein color should become more pronounced. Ask students to propose explanations for the behavior of the conductivity light bulb.

Demonstration 5: pH and Environmental Gases

Dissolved gases often produce acidic or basic solutions.

Materials

5 Flasks, conical or round, 250-mL
Universal indicator solution
Gas sources for CO_2, NH_3, and SO_2
Dry ice, if available

Safety

Dry ice can cause severe frostbite and should be handled with gloves or tongs. Other gas sources need to be handled carefully. Use the fume hood. Goggles should be worn.

Directions

Add 100 mL of distilled water to each flask. Add universal indicator solution to produce an intense color. Bubble each individual gas, one at a time, into the distilled water in the separate flasks. Add a piece of dry ice to the fourth flask. The fifth flask can be left with distilled water as a control. Have students observe and interpret the results.

Demonstration 6: pH Measurement

pH can be measured in several convenient ways.

Materials

Narrow and broad range pH paper
pH pen or pH meter
Solutions with various pH values

Safety

Become familiar with the operating instructions for the pH measuring device.

Directions

If a pH meter is available and used, show students not only how to use the meter but also how to calibrate the instrument at a pH close to the value to be measured. Adapt this demonstration to your needs and available pH measuring devices. Indicators may also be used to supplement the pH paper or meter observations.

Demonstration 7: Silent Demo—Reactivity and Concentration

As the concentration of an acid decreases, reactivity with metals decreases also.

Materials

6 M Hydrochloric acid, HCl
Three samples (about 4 cm) of magnesium metal ribbon
An overhead projector or a light box

Safety

6 M HCl is corrosive and must be handled with care. Goggles should be worn.

Directions

Inform the students that this will be a silent demonstration for both the teacher and students. Information can, however, be written on the board. Place one piece of Mg ribbon into 50 mL HCl in a 250-mL beaker. With a large-face timer or stop-watch, indicate the time needed for the Mg to react completely on the board. Show the students a second beaker with 50 mL water; then pour the acid into the water beaker. Again note the time required for a complete reaction. The teacher can continue the dilutions as long as there is an observable reaction. The discussion should contain other examples where concentration is a big factor in reaction rates and toxicity. Hopefully the students will begin to understand the importance of developing a quantitative method of expressing concentration, which will follow in the lesson.

Suggestions for Other Demonstrations

Bo, Sk, Su, and Sq refer to monographs by Borgford, Shakhashiri, Summerlin, and Sarquis listed in *References*. *CAUTION: Check each reference for appropriate safety guidelines before performing these demonstrations.*

Natural indicators. Extractions of colored dyes from plants that can function as acid-base indicators. *[Sk v3 50; Sq CF 27; Bo CA 92]*

Properties of acids and bases. Demonstrates classical properties of acids and bases. *[Sk v3 58; Su v2 171]*

pH of familiar products. pH measurement of household chemicals using indicator solutions or a pH meter reveals that many substances are acids and bases. *[Sk v3 65; Bo CA 133]*

Disappearing ink. A blue "ink" spot on white cloth is caused to fade and disappear. *[Su v2 174; Sq CF 325]*

Invisible painting. A pattern is painted on a sheet of paper with phenolphthalein (colorless), and the pattern is developed when the paper is sprayed with sodium hydroxide solution. *[Sk v3 47]*

Cylinders of color. Create color changes in solutions containing several different acid-base indicators. *[Sk v3 114 (uses dry ice); Sq CF 127 (uses Alka Seltzer™), Su v1 38]*

Acid-Base nature of common gases. The properties of several gases are revealed: carbon dioxide in breath *[Bo CA 15]*, sulfur dioxide *[Su v2 165]*; and ammonia *[Sq CF 59]*.

Sealed-bag reactions with acids and bases. A reaction occurs among three substances mixed in a sealed plastic bag *[Sq CF 345]* or between two substances in a sealed plastic bag *[Sk v3 100]*.

Acid-base titration/conductivity. Conductometric titrations of barium hydroxide and sulfuric acid. *[Sk v3 152]*

Making hydrogen gas from an acid and a base. Hydrogen gas is collected in a balloon attached to each of two flasks. One flask contains aluminum foil and hydrochloric acid; the other, aluminum foil and sodium hydroxide. *[Su, v2, 33]*

Acid-base indicators: A voice-activated chemical reaction. Allow students to speak into a flask containing bromthymol indicator. After several tries, the solution will turn yellow. *[Su v1 40]*

Miscellaneous applications. Measuring acid rain *[Su CA 191]*, measuring soil pH *[Bo CA 194]*, what's in baking powder *[Bo CA 223]*, carbonates in rocks, minerals, and eggshells *[Bo CA 17; Sq CF 305]*.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What are common observable properties of acids and bases? *[Acids taste sour, cause blue litmus to turn red, liberate hydrogen gas when they react with certain metals, and neutralize bases. Bases taste bitter, feel slippery, cause red litmus to turn blue, and neutralize acids.]*
2. Identify acids and bases you have encountered in and around your home. *[Baking soda (sodium bicarbonate, $NaHCO_3$, base), drain cleaner (sodium hydroxide, $NaOH$, base), vinegar (acetic acid, $HC_2H_3O_2$, acid), citrus fruits (citric acid, acid), battery acid (sulfuric acid, H_2SO_4, acid), antacids (e.g., Milk of Magnesia, $Mg(OH)_2$, base), household ammonia (aqueous ammonia, NH_3, base)].*
3. How does the concentration of hydrogen ion vary as pH changes? *[As the pH increases, the concentration of H^+ decreases.]*
4. From what acid and base could we synthesize sodium nitrate, $NaNO_3$? *[The anion of the salt comes from the base and the cation comes from the acid. For $NaNO_3$, the sodium ion, Na^+, is supplied by sodium hydroxide, $NaOH$, and the nitrate ion, NO_3^-, is supplied by nitric acid, HNO_3.]*
5. What are some problems associated with "acid rain"? *[Acid rain is a dilute solution of sulfuric acid and nitric acid. Because acids are corrosive, acid rain can cause damage to humans, metal and brick buildings, marble and limestone structures, and the environment (kill trees and crops, change pH of soil,* etc. *See "Checking Out Acids and Bases" in* ChemMatters, *April 1983.]*
6. Why do some shampoo bottles say "acid-balanced"? *[These shampoos claim to maintain the normal acid pH of hair and contain an acid, usually citric acid, to counteract the alkalinity of the detergents. See "Checking Out Acids and Bases" in* ChemMatters, *April, 1983.]*
7. What do antacids do? Are there any problems with taking too much? *[Antacids neutralize excess stomach acid brought on by stress, overeating, or disease. Some antacids contain sodium ion (Alka Seltzer™), which can aggravate hypertension. Milk of Magnesia, $Mg(OH)_2$, in large doses has a laxative effect. See "Checking Out Acids and Bases" in* ChemMatters, *April 1983.]*

Counterintuitive Examples and Discrepant Events

1. Diluting a weak electrolyte ($HC_2H_3O_2$) with water increases the electrical conductivity (see *Suggested Laboratory Activity*, Strong *vs.* Weak Acids).
2. Carbon dioxide bubbled through limewater causes a precipitate to form. Continued bubbling causes the precipitate to disappear. [See Underground sculpture. (1984). *ChemMatters, 1*(2), 10-11.]
3. Some active metals react with either acids or bases to produce hydrogen gas. (For example, aluminum will react with either hydrochloric acid or sodium hydroxide, releasing hydrogen gas. See *Suggestions for Other Demonstrations*, Making hydrogen gas from an acid and a base.)
4. The same amount of hydrogen gas will be produced when a sample of an active metal is added to equal volumes of concentrated acid solution and dilute acid solution (*if* the metal is the limiting reactant).
5. Bicarbonate salts can be used to neutralize either an acid or a base. This can be illustrated by the following equations:

$$HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$$

$$HCO_3^-(aq) + OH^-(aq) \rightarrow H_2O(l) + CO_3^{2-}(aq)$$

Metaphors and Analogies

1. $[H^+]$ *vs.* pH: A see-saw relationship (see *Pictures in the Mind*); one goes up (increases) while the other goes down (decreases).
2. A proton shifting from an acid to a base can be likened to a baseball being thrown from a pitcher (the acid) to a catcher (the base).
3. Universal indicator color changes follow the colors in the rainbow as the pH moves from 2 to 10. The name ROY G BIV helps keep the colors straight: Red, Orange, Yellow, Green, Blue, Indigo, Violet. A pH of 7 produces a yellow-green hue.

Pictures in the Mind

1. **Ionization**. Graphical pictorial representation of the behavior of acids of different strengths in aqueous solution.
   a. A strong acid is completely dissociated.
   b. In a weak acid only a small fraction of molecules is dissociated in solution.

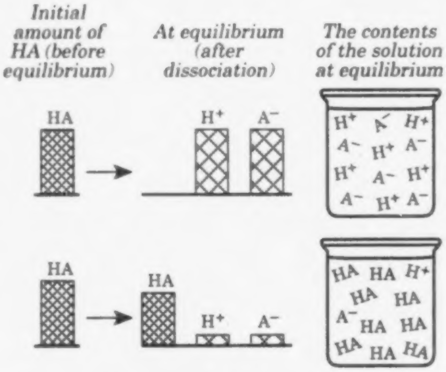
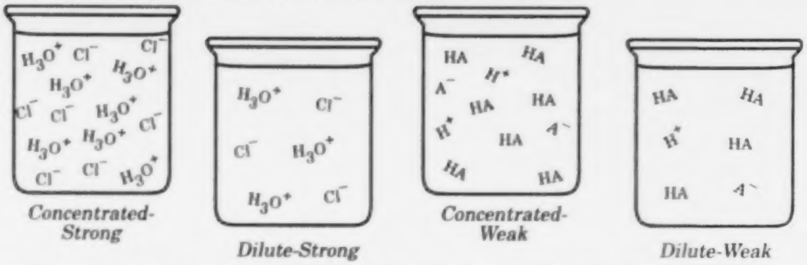

Initial amount of HA (before equilibrium)
At equilibrium (after dissociation)
The contents of the solution at equilibrium
HA
H+
A−
H+ A− H+
A− H+ A−
H+ A− H+
A− H+ A−
HA
HA
H+
A−
HA HA H+
HA HA
A− HA HA
HA HA HA

Figure 11. Picturing ionization of strong acids (upper) and weak acids (lower).

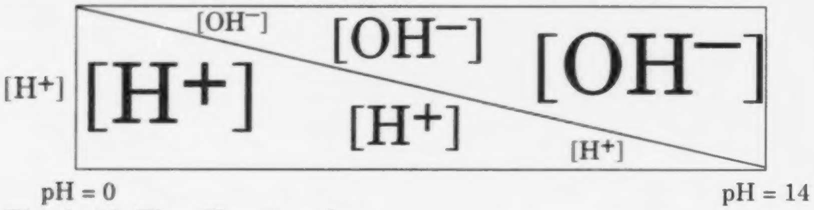
2. Pictures of the composition of a concentrated-strong acid, a dilute-strong acid, concentrated-weak acid, and a dilute-weak acid.



H_3O^+
Cl^-
HA
H^+
A^-
Concentrated-Strong
Dilute-Strong
Concentrated-Weak
Dilute-Weak

Figure 12. Picturing acid strength and concentration.

3. pH rectangle



[H+]
[H+]
[OH−]
[OH−]
[OH−]
[H+]
[H+]
pH = 0
pH = 14

Figure 13. The pH rectangle.

4. [H+] versus [OH−]: A see-saw

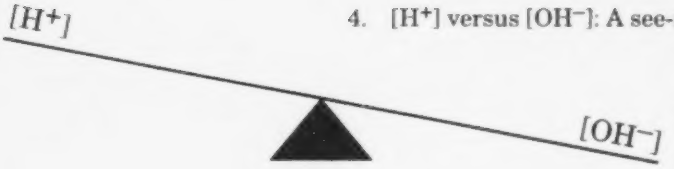


Figure 14, See-saw demonstrates inverse relationship of [H+] and [OH−].

5. pH versus [H+]: a see-saw

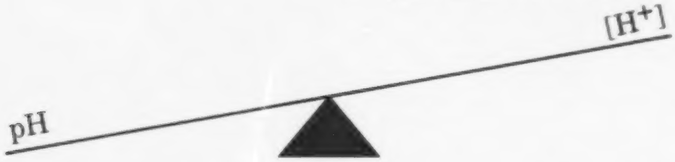
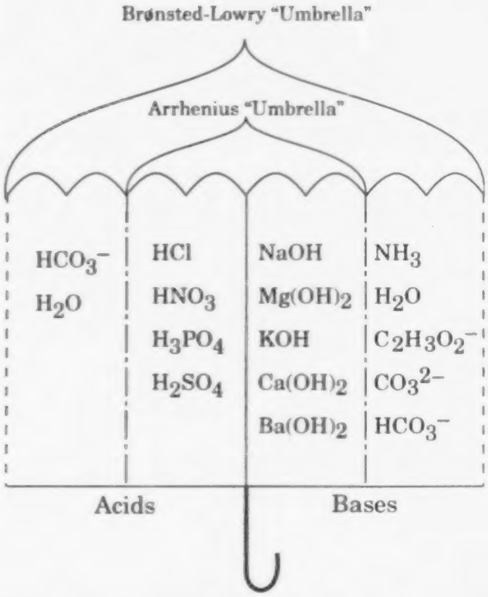


Figure 15. A pH/[H+] see-saw.

6. Bronsted-Lowry "Umbrella"



Brønsted-Lowry "Umbrella"
Arrhenius "Umbrella"
HCO_3^-
H_2O
HCl
HNO_3
H_3PO_4
H_2SO_4
NaOH
$Mg(OH)_2$
KOH
$Ca(OH)_2$
$Ba(OH)_2$
NH_3
H_2O
$C_2H_3O_2^-$
CO_3^{2-}
HCO_3^-
Acids
Bases

Figure 16. Umbrellas of acid-base definitions.

Language of Chemistry

Brønsted-Lowry acid proton donor.

Brønsted-Lowry base proton acceptor.

concentration amount of solute in some specified amount of solution.

end point point in a titration when the indicator changes color.

indicator weak acid or base that changes colors over a narrow pH range.

Lewis acid electron pair acceptor.

Lewis base electron pair donor.

molarity concentration term defined as the number of moles of solute per liter of solution.

neutralization reaction between an acid and a base in which the acid and base properties disappear.

strength percentage of ionized molecules of acid (or base).

TIPS FOR THE TEACHER

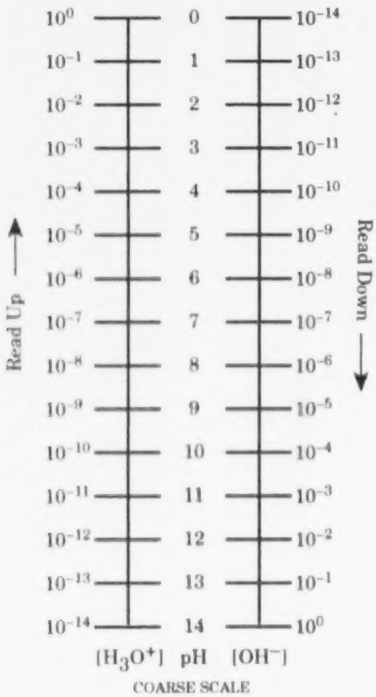


Figure 17. Scales for pH-concentration relationships.

1. The word acid comes from the Latin *acidus*, for sour. One property of acids is their sour taste, such as that experienced with yogurt, pickles, and lemons.
2. The terms strong and weak are not synonymous with concentrated and dilute, respectively (see *Common Misconceptions*).
3. For consecutive integer values of pH, H^+ concentrations differ by a factor of 10. For example, solutions of pH = 1 ($[H^+] = 10^{-1}$ M) and pH = 2 ($[H^+] = 10^{-2}$ M) have a H^+ concentration ratio of 10, that is, $10^{-1}/10^{-2} = 10$.
4. Solutions with high hydrogen ion concentration have a low pH. A solution with pH = 1 has a $[H^+] = 10^{-1}$ M; a solution with pH = 6 has a $[H^+] = 10^{-6}$ M. Figure 17 provides convenient conversion scales between pH and concentration.
5. pH-pOH square. Equations over or next to arrows are conversion equations. Given one corner, the other can be calculated using the conversion equation.

$$\begin{tikzcd}
{[H^+]} \arrow[r, leftrightarrow, "pH = -\log[H^+]"] \arrow[d, leftrightarrow, "{[H^+][OH^-] = 10^{-14}}"'] & pH \arrow[d, leftrightarrow, "pH + pOH = 14"] \\
{[OH^-]} \arrow[r, leftrightarrow, "pOH = -\log[OH^-]"'] & pOH
\end{tikzcd}$$

Figure 18. pH-pOH square.

6. A pH rhyme from Kolb, D. (1979). *Journal of Chemical Education*, *56*, 53:

> For coffee it's 5, for tomatoes it's 4;
> While household ammonia is 11 or more.
> It's 7 for water, if in a pure state,
> But rain water is 6, and sea water is 8.
> It's basic at 10, quite acidic at 2,
> And well above 7 when litmus turns blue.
> Some find it a puzzlement. Doubtless their fog
> Has something to do with that negative log.

7.

Solution	pH
Battery acid	0
Stomach acid	1.4 - 1.8
Lemon juice	2.1
Vinegar	2.9
Soda water	3
Wine	3.5
Tomato juice	4
Black coffee	5
Urine, sour milk	6
Rain water	6.5
Pure water (25 °C)	7
Blood	7.35 - 7.45
Sea water	8
Soaps, shampoos	8 - 9
Detergents	9 - 10
Milk of magnesia	10
Household ammonia	11.9
Liquid bleach	12
Household lye	14

Figure 19. Approximate pH of some common solutions.

8. Remember the color litmus becomes in acidic and basic solutions:

Blue Base
reD aciD

or

aci
re D

ase
Blue

9. Generally, concentrations of acidic and basic solutions are expressed as moles per liter (M). Commercial acids and bases are sold as percent by mass. For example, concentrated hydrochloric acid is 37% by mass. The concentration of sulfuric acid in rain water is measured in parts per million (ppm). (See *Introductory* module for conversion table.)

10. Practice writing balanced equations for neutralization reactions.
11. Review safety precautions when handling acids and bases.

Common Student Misconceptions

1. **"Substances containing H are acidic; substances containing OH are basic."**

   Many substances that contain H are not acids and many substances that contain OH are not bases. Table sugar (sucrose), $C_6H_{12}O_6$, contains H and OH; however, when dissolved in water it dissolves as intact molecules and does not ionize to produce any H^+ or OH^- ions, as shown by the equation: $C_6H_{12}O_6(s) + H_2O(l) \rightarrow C_6H_{12}O_6(aq)$. Since sugar molecules do not lose any H^+ or OH^- ions, sugar is not considered an acid. Alcohols have a characteristic hydroxyl group, OH, covalently bonded to the rest of the molecule. When alcohols dissolve in water they also dissolve molecularly and do not ionize. Since no OH^- ions are released, alcohols are not bases—they actually are weak acids! Another approach is to consider the following substances containing OH groups: NaOH, CH_3OH, CH_3COOH, and $SO_2(OH)_2$. The first is a strong base in water, the second (methanol) has neither acidic or basic properties in water, the third (acetic acid) is a weak acid in water, and the latter (H_2SO_4) is a strong acid in water. The concept of bonding (see *Chemical Bonding* module) is useful with this approach.

2. **"When a 'proton donor' acid reacts, the nucleus of an atom loses a proton."**

   When we speak of acids as proton donors we are talking about the single proton in a hydrogen ion, H^+, not a proton from the nucleus of some other atom. Although chemists commonly refer to acids and bases as proton donors and proton acceptors, it is important to realize that H^+ is being moved from one chemical species to another, not a proton from one nucleus to another. Furthermore, if nuclear protons *were* transferred, then elements would change their identities in acid-base reactions, since the atomic numbers of the atoms involved would be altered. Another effective way to deal with this confusion is to have students draw a "structure" for the hydrogen atom. Attention should be drawn to what is in the hydrogen nucleus, and students ask, "If we remove an electron from a hydrogen atom, what is left?" It is then fairly easy for students to understand that a hydrogen ion is a bare proton, but not the proton drawn from the nucleus of just any atom.

3. **"Strength and concentration mean the same thing."**

   These two terms are often confused by students. Concentration refers to the amount of solute in some specified amount of solution. Strength refers to the percent of molecules that ionize and form ions in solution. Hydrochloric acid is considered a strong acid because in aqueous solutions nearly all the molecules ionize to produce H^+ and Cl^- ions. Acetic acid is considered a weak acid because only about 1% of the molecules normally ionize. It is interesting to note that the percent of acid molecules that normally ionize in a weak acid actually increases as the acid concentration is decreased through dilution with water. Thus, acetic acid actually becomes slightly stronger as the solution is diluted.

4. **"Concentration and amount mean the same thing."**

Some students have difficulty distinguishing the concepts of amount and concentration. By international convention (SI), the *amount of a substance* is measured in *moles*. Solution concentration is frequently expressed as Molarity (M), which is a ratio of moles of solute to liters of solution. Thus, simply stating Molarity does not give amount of substance; the volume must also be known. The amount of acid, in moles, is found from the relationship:

Moles = Molarity X Volume (in liters)

Thus, the amount of solute in these solutions: 1.0 L of 2.0 M
2.0 L of 1.0 M
0.50 L of 4.0 M

is 2.0 mol in each. It is apparent that the concentrations, 2.0, 1.0, and 4.0 M, are not identical. Concentration is not, then, the same as amount. This is emphatically illustrated by discussing the relative effects of 1 drop of concentrated sulfuric acid (0.05 mL of 18 M H_2SO_4) and 9 L of 0.1 M sulfuric acid. While both contain 0.9 moles of acid molecules, a single drop of the former solution causes a painful burn, while the latter is harmless. Demonstrate this using a paper towel rather than skin! The concentrated acid eats through the paper, leaving a carbonized patch; but dropping a towel into a plastic bucket of the diluted acid (1 drop in 9 L) causes no change.

You may choose to illustrate concentration *vs.* amount using the following visual system, which does not refer to acid-base concepts, but uses food coloring (see *SourceView:* Problem Solving, also). Place 3-4 drops of food coloring in a 100 to 250-mL graduated cylinder filled about a third full with water. Stir. Set this on an illuminated overhead projector stage. Add water, having students observe from top and side. Ask if any more dye was added. Point out that the side view shows the dilution effect while the top view (color constant) indicates the total amount of dye substance. For an explanation of how color and concentration are related, read on.

The color varies according to Beer's Law.

$A = abc$ where,

A = Absorption
a = Molar absorptivity (a constant for a given colored species)
b = Light path length or solution "thickness," and
c = Concentration of colored species

As water was added, a remained constant, c decreased, cross section b remained constant. Thus from the side, $A = abc$ decreased. From above, however, b increased as much as c decreased so $A = abc$ was constant.

5. **"Ammonium hydroxide is the name of the solution formed when ammonia gas dissolves in water."**

Ammonia gas does readily dissolve in water. After it has dissolved, the equilibrium represented by the following expression is eventually attained.

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq).$$

Sometimes this solution is called ammonium hydroxide. However, the equilibrium lies far to the left; few ammonium ions, NH_4^+, or hydroxide ions, OH^-, are present. The compound NH_4OH has not been isolated. Water solutions of ammonia are weak electrolytes. Aqueous ammonia is a better name. Ammonia is a weak base.

6. **"A neutral solution always has a pH equal to 7."**

   The percent of water molecules that ionizes depends a great deal on temperature. Only at 25 °C does $[H^+] \times [OH^-] = 10^{-14}$, giving a pH of 7 at neutrality. As temperature increases so does the percent of water molecules that ionize. Thus, at higher temperatures the pH of neutrality decreases. At body temperature, 37 °C, the pH of neutral water is 6.81.

7. **"A neutralization reaction produces a neutral solution."**

   A neutralization reaction is one in which acid and base properties disappear when an acid solution is added to a base solution. A solution of an ionic compound (a salt) results. This solution may or may not be "neutral' (that is, have a pH of 7; see the Hydrolysis section under *Extensions*).

8. **"A pH = 0 means that no acid or base is present."**

   A solution with a pH = 0 would be *very* acidic—in fact, the solution would have a hydrogen ion concentration of 1 M. [pH = –log(1) = 0]

9. **"The compound HNaO would be considered an acid."**

   Students often think that the way a compound's formula is written determines whether the compound is an acid or a base. Because of the way that Na, O, and H interact, they produce a compound commonly known as sodium hydroxide. No matter how we write the formula (HNaO or NaOH), the same basic compound is produced. A similar confusion can arise when acetic acid is written as $HC_2H_3O_2$ or CH_3COOH. In both cases an acidic compound is represented. The difference is that the first is written to correspond to the way most mineral acids are written, while the second gives a better indication of its actual organic carboxylic structure. It's not how we write the formulas, but how the atoms in the structure actually interact that determines the properties of compounds.

10. **"The end point and neutralization point in a titration are the same."**

    Not so—neutralization involves consuming all acid and base; end point depends on indicator color change.

11. **"A concentrated acid is a pure substance."**

    Acids are usually used in aqueous solutions and are mixtures, not pure substances.

12. **"Concentrated acids are much more dangerous than concentrated bases."**

    Most beginning chemistry students have some knowledge of the dangers of concentrated acids. (Battery acid, horror movies, *etc.*) These same students have little or no knowledge of bases and their properties. Students should be reminded that most "bowl cleaners" are concentrated bases because they are good at dissolving organic material. When students realize they are made up of organic material, they begin to build a healthy respect for bases.

13. **"Only pH values of 1-14 exist."**

    The definition of pH ($-\log[H^+]$) permits values of pH outside the 1-14 range, although they are seldom encountered. Substituting molarity of commercial concentrated hydrochloric acid (12 M) in the definition equation gives a pH = –1.1; for 18 M sulfuric acid, the pH = –1.3. Similarly, the 8 M sodium hydroxide solution used to make soap from animal fat has $[H^+] = 1.25 \times 10^{-15}$ (that is, $10^{14}/8$), and a corresponding pH = 14.9.

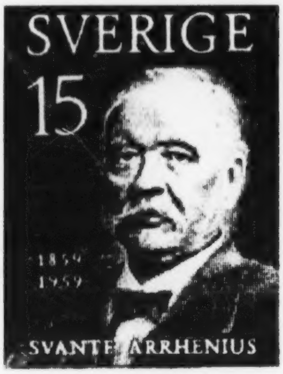
    Solutions of such extreme concentrations are seldom used, except as concentrates for making the dilute solutions of pH 1-14.

14. **"For sulfuric acid, to say that dissociation is complete means the acid has lost both protons."**

When we say that dissociation of sulfuric acid is complete, we are not saying that the species in solution are two hydrogen ions and one sulfate ion for each molecule dissolved. The fact is, sulfuric acid is "strong" only in regard to removal of the first hydrogen ion.

HISTORY: ON THE HUMAN SIDE

1. **Svante August Arrhenius** taught himself to read at age three and graduated from high school as the youngest and brightest in his class. As a student at the University of Uppsala he studied how electricity passed through solutions. He proposed that the resulting electric current from an aqueous sodium chloride solution was carried not by ordinary sodium and chlorine atoms, but by atoms carrying an electric charge—sodium ions and chloride ions. These ideas, which Arrhenius presented in 1884 in his Ph.D thesis, met with resistance. He was awarded the lowest possible passing grade by his examiners. Fortunately, Van't Hoff and Ostwald, two influential physical chemists, were impressed and encouraged him to continue work in physical chemistry. In 1887, Arrhenius proposed that the characteristic properties of acids in water solution are the properties of hydrogen ion and those of bases, the properties of hydroxide ion. When J. J. Thomson discovered the electron in the 1890s, the idea of "ions" that Arrehenius proposed in his thesis suddenly became credible. In 1903, for the same thesis that had barely earned him a passing grade in his Ph.D examination, Arrhenius was awarded the Nobel Prize in chemistry. In 1959 Sweden issued a postage stamp on the occasion of his birth centenary.



2. **Johannes Brønsted** was born in Denmark in 1879. When he was 13 his father died; he was to follow in his father's profession of civil engineer. However, he was interested in chemistry and studied it in college. In 1908 he earned his doctorate and was selected as a new professor of chemistry at the University of Copenhagen. While studying how acids and bases catalyzed reactions in 1921, he clarified what acids and bases were. Since the properties of acids and bases were opposites, Brønsted felt it made more sense to supply definitions that were opposites. In 1923 he suggested that if acids were species that gave up a proton, bases are species that accept a proton. At the same time, Thomas Lowry, a British chemist, independently proposed the same definition. This definition represented a greater flexibility that extended acid-base notions into areas in which the old view was inadequate. For examples, ions could be classified as acids or bases and acid-base reactions could occur without water being present.

HUMOR: ON THE FUN SIDE

1. Confucius: "Chemist who falls in acid is absorbed in his work!" (Chem 13 News, January 1975, p. 838.)
2. Rhyme:

Johnny was a chemist
Now Johnny is no more
For what he thought was H two O
Was H two S O four!

3. Student responses on exams:
    - a. A strong base is a salt that produces a strong acid when reacted with water.
    - b. A strong acid is one that will not make a light burn bright.
    - c. Filter paper is an indicator.
4. Sign on a bumper sticker: *Chemistry is pHun!*
5. Rhyme:

Mamie often wondered why
acids trouble alkali.
Mamie in a manner placid
fed the cat some boric acid;
whereupon the cat grew frantic
executing many an antic.
Concluded Mamie from its pace,
"Pussy cat must be a base."

(*Chem 13 News, November 1988, p. 13.*)

6. Word Search (see *Appendix* for master copy)

Z S N M A T K B Z ~~R U O S~~ T M P Y
X T S Q R A D H V K Q K Y O S E W
R P H T D R B U X E R I L D U R H
E U Y O R U M X J O S A I Q H B Y
T R D F H E P G N H R W N A V W D
T M R Z D V N O Q I O P O Y K M R
I R O C V M D G T R T T I X D O O
B I G V F N J Y T P A V W L X D X
A U E A O X Y N W H C C R Z J S I
M Y N T K A T H ~~T N I O P D N E~~ D
G Y O P V T E I D Z D F E S I K E
M R O Y D I A V C M N G J T H S J
P Q Z ~~N E U T R A L I Z E~~ I N I R

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Weak acids or bases that change color over a narrow pH range.
2. To eliminate the acidic or basic character of a solution.
3. Term used to denote the degree of ionization of an acid or base.
4. The point in a titration when the indicator changes color (2 words).
5. Concentration of a solution defined in terms of moles of solute per liter of solution.
6. Ion most often associated with an acid.
7. Ion most often associated with a base.
8. Taste most often associated with acids.
9. Taste most often associated with bases.
10. Brønsted-Lowry acid (2 words).

Answers: 1. INDICATORS 2. NEUTRALIZE 3. STRENGTH 4. END POINT 5. MOLARITY 6. HYDROGEN 7. HYDROXIDE 8. SOUR 9. BITTER 10. PROTON DONOR

7. See relevant cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 16: The Proton in Chemistry," World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
2. CHEM Study films/videos available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.
   - a. *Acid-Base Indicators*
   - b. *Nitric Acid*
3. Moore, J. (Ed.). (1988). Acid-base indicator simulation for Apple II microcomputer. *Journal of Chemical Education, 65,* 798. Computer series.
4. Atkinson, G., Doadt, E., and Rell, C. (1986). Master variable diagrams for acid-base systems for an IBM personal computer. *Journal of Chemical Education, 63,* 841.
5. Software published by *JCE: Software*, a publication of the Journal of Chemical Education, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   - a. *The Acid-Base Package,* by Richard Ramette. Vol. II B, Bo. 2, for IBM PS/2 PC-compatible computers.
   - b. *Acid-Base Package,* by Richard Ramette and Jon L. Holmes. Vol. IV C, No. 2, for the Apple Macintosh.
6. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
   - a. For the Apple II computer: AP 501
   - b. For IBM PCs and PC-compatibles: PC 2701, PC 3901
7. Videodisc published by *JCE: Software,* a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Acid Rain and Limestone" and "Lime Water and Carbon Dioxide," two chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
8. *Acids, Bases and Salts*, a 20 min. video available from Coronet/MTI Film and Video, 108 Wilmot Road, Deerfield, IL, 60015; (800) 621-2131; (708) 940-3640 (FAX).

INSTRUMENTATION

1. Vernier's Voltage Plotter with pH electrode (Vernier Software Science Supply, 2920-T.S.W. 89th Street, Portland, OR 97225).
2. Nester model and Corning model pH meters (Flinn Scientific, Inc. Chemical Catalog/Reference Manual, 131 Flinn Street, P. O. Box 219, Batavia, IL 60510-0219).
3. Corning Conductivity Meter (Flinn Scientific).
4. Conductivity Apparatus: Russo, T. (1986). A low-cost conductivity apparatus. *J. Chem. Ed.*, *63*, 981; Gadek, F. (1987). Easily made electronic device for conductivity experiments. *J. Chem. Ed.*, *64*, 628; Gadek, F. (1987). A commercially available device for conductivity experiments. *J. Chem. Ed.*, *64*, 281; and Vitz, E. (1987). Conductivity of solutions apparatus. *J. Chem. Ed.*, *64*, 550.
5. Oyster-Shell Type pH Meter. Extech, 150 Bear Road, Waltham, MA 02154, 617-890-7440. This instrument costs less than $200. It has two calibration adjustments to give results comparable to more expensive pH meters.

Links/Connections

WITHIN CHEMISTRY

1. **Periodicity**

Metal oxides usually form basic solutions in water; aqueous solutions of nonmetal oxides are acidic.

$$\underset{\text{Metal oxide}}{K_2O} + H_2O \rightarrow 2KOH$$

$$\underset{\text{Nonmetal oxide}}{SO_3} + H_2O \rightarrow H_2SO_4$$

The acidity of the hydrogen halides increases in the sequence

$$HF < HCl < HBr < HI$$

The basicity of Period 3 hydroxides decreases in the sequence

$$NaOH > Mg(OH)_2 > Al(OH)_3$$

The acidity of Period 3 oxyacids increases in the sequence

$$H_3PO_4 < H_2SO_4 < HClO_4$$

2. **Solutions/solubility**

Most acids tend to be water soluble. NaOH and KOH are quite soluble in water. $Ca(OH)_2$ is moderately water soluble. Acidic solutions dissolve some metals (Fe, Zn, and Al), releasing hydrogen gas. Zinc and aluminum will dissolve in base. Drano™ consists of solid NaOH and Al turnings. When water is added to the mixture, hydrogen gas is evolved.

$$2Al(s) + 2NaOH(aq) + 2H_2O(l) \rightarrow 2NaAlO_2(s) + 3H_2(g)$$

3. **Equilibrium; Ionization Constant**

The dissociation of a strong acid or base is complete. The ionization constant expression is a measure of the equilibrium for the ionization of a weak acid or base. For example, the equilibrium established when the weak acid, acetic acid, ionizes,

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

is given by the ionization constant expression:

$$K_a = K[H_2O] = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

Water has the important characteristic of being able to act as either an acid or base. This tendency allows two water molecules to react accordingly

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

The equilibrium constant expression for the process is

$$K_w = K[H_2O][H_2O] = [H_3O^+][OH^-]$$

K_w, the ion product constant for water, has a value of 1.0×10^{-14} at 25 °C and 2.4×10^{-14} at 37 °C (body temperature). This is significant because it means that although at 25 °C a neutral solution has a pH = 7, at 37 °C, the pH of a neutral solution is less than 7 (6.81).

4. **Electrochemistry**

Electrolysis of water produces an acidic solution at the anode and a basic solution at the cathode. When an electrical current is passed into graphite electrodes immersed in a sodium sulfate solution, hydrogen gas and hydroxide ions are produced at the cathode:

$$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$$

At the anode, oxygen gas, and hydrogen ions are formed:

$$2H_2O(l) \rightarrow O_2(g) + 4H^+(aq) + 4e^-$$

The OH^- and H^+ ions combine to reform water. The overall reaction is

$$2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$$

5. **Stoichiometry**

See the *Stoichiometry* module.

6. **Bonding**

Hydrogen halides have a polar covalent bond. $\mathrm{H} : \underset{\cdot\cdot}{\overset{\cdot\cdot}{\mathrm{F}}} :$ $\overset{\delta+\ \delta-}{\mathrm{H{-}F}}$

In oxyacids, hydrogen is bonded to oxygen. The central atom becomes very important in determining acidity.

H_2SO_4 is H-O-SO_2-O-H and $HClO_4$ is H-O-ClO_3

Chlorine is more electronegative than sulfur and is able to weaken the O-H bond, increasing the acidity of $HClO_4$ over that of H_2SO_4.

7. **Thermochemistry**

Thermochemistry is a study of the quantity of heat absorbed or evolved by chemical reactions. The neutralization of a strong acid and strong base, represented by the equation

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l) + \text{Heat}$$

is highly exothermic. Acid or base spills on the body should not be neutralized since the affected area could actually experience a heat burn due to high heat of neutralization. Instead, dilute the acid or base by flushing with lots of water.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Environmental Science**

Acid rain is rain with a pH less than 5.6. Normal rain water has a pH of 5.6 due to dissolved carbon dioxide, CO_2. Acid rain is primarily a dilute mixture of sulfuric acid and nitric acid that is formed when sulfur and nitrogen oxides (nonmetal oxides) dissolve in rain or snow. Acid rain destroys lakes, kills forests and crumbles buildings and statues.

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Acid mine water is due to sulfuric acid produced from iron sulfide (pyrite, FeS_2) in coal. Acid mine water is associated with strip mining, produced when pyrite deposits are exposed to the atmosphere.

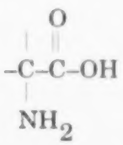
$$2FeS_2(s) + 7O_2(g) + 2H_2O(l) \rightarrow 2FeSO_4 + 2H_2SO_4$$

$Ca(OH)_2$ and $Al(OH)_3$ are important in water purification. During the settling stage, $Ca(OH)_2$ and $Al_2(SO_4)_3$ are added to produce $Al(OH)_3$, a sticky, gelatinous precipitate that settles out slowly, carrying suspended dirt particles and bacteria with it.

2. Biochemistry

Gastric (stomach) acid is essentially hydrochloric acid with a pH of about 1.6-1.8. It activates protein-digesting enzymes, which are responsible for protein digestion in the stomach.

Proteins are polymers of amino acids. Amino acids have two functional groups, the basic amino group, NH_2, and the acidic carboxyl group, COOH:

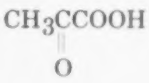


$$-\underset{\displaystyle NH_2}{\underset{|}{\overset{|}{C}}}-\overset{\displaystyle O}{\overset{\|}{C}}-OH$$

About 20 different amino acids make up the proteins in the body.

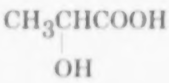
DNA (deoxyribonucleic acid) and RNA (ribonucleic acid) are complex molecules whose biological activity is dependent on the composition of organic nitrogen bases present in their structures. DNA is the essential hereditary material of all living cells. RNA controls the production of proteins, the building blocks of living cells.

Physical activity is a series of muscle contractions, each requiring energy. Muscles store energy as glycogen. This glycogen is the storage form of glucose in the blood arising from carbohydrate in the diet (candy bar, fruit, potatoes, pancake syrup, *etc.*). When muscles contract, the glycogen is converted to pyruvic acid and, if sufficient oxygen is present, to CO_2 and H_2O. This is aerobic exercise or aerobics. If sufficient oxygen is unavailable, pyruvic acid is converted to lactic acid. This is anaerobic exercise. If the buildup of lactic acid continues, the muscles are less responsive to contraction, and the person experiences some fatigue and even pain. After exercising, more oxygen becomes available to convert the lactic acid to pyruvic acid, and the latter is converted to CO_2, H_2O, and energy. Sprinting and weight lifting are typical anaerobic activities; a marathon race is largely aerobic. Warm-up activities before exercise help to prevent rapid buildup of lactic acid in the muscle tissues.



$$CH_3\underset{\displaystyle O}{\underset{\|}{C}}COOH$$

lactic acid



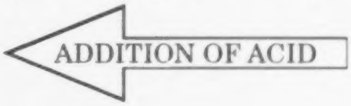
$$CH_3\underset{\displaystyle OH}{\underset{|}{C}}HCOOH$$

pyruvic acid

To insure necessary biological functions, body fluids are maintained within a narrow pH range. For example, protein molecules within cells can be altered if the pH changes too much in either direction of the optimum pH. This alteration can change the usual properties of the protein such that the protein would not be able to carry out its normal function.

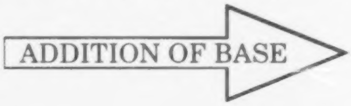
The H_2CO_3/HCO_3^- buffer is the major buffer in blood. This buffer counteracts either added acid or bases. Acidic by-products (H^+) of metabolic cycles are neutralized by $HCO_3^-(aq)$ forming $H_2CO_3(aq)$. Excess $H_2CO_3(aq)$ is removed from the body as $CO_2(g)$ in the lungs.

$$CO_2\,(g) + H_2O \rightleftharpoons H_2CO_3\,(aq) \rightleftharpoons HCO_3^-\,(aq) + H^+\,(aq)$$



Excess base is neutralized by $H_2CO_3(aq)$ forming $HCO_3^-(aq)$. To compensate for the diminished $H_2CO_3(aq)$, $CO_2(g)$ dissolves in water to form carbonic acid.

$$CO_2\,(g) + H_2O \rightleftharpoons H_2CO_3\,(aq) \rightleftharpoons HCO_3^-\,(aq) + H^+\,(aq)$$



3. **Geochemistry**

Magnesium, Mg, is the lightest structural metal in common use. It is obtained from sea water. Lime, CaO, is added to sea water to precipitate magnesium hydroxide, $Mg(OH)_2$. Magnesium hydroxide is then filtered and neutralized with hydrochloric acid. After evaporation of the water, the resulting $MgCl_2$ is melted and electrolyzed to give magnesium metal.

After recovery from coal and petroleum or even sulfur mines, sulfur is converted to sulfuric acid in four steps: (1) the sulfur is burned in air to sulfur dioxide, SO_2; (2) sulfur dioxide is passed over a hot platinum surface where it is converted to sulfur trioxide, SO_3; (3) sulfur trioxide is passed into sulfuric acid where pyrosulfuric acid, $H_2S_2O_7$, is formed, and (4) the latter is diluted with water to give sulfuric acid, H_2SO_4. Sulfuric acid is the number one chemical substance in terms of total mass produced by U.S. industry (see *Industrial Inorganic Chemistry* module).

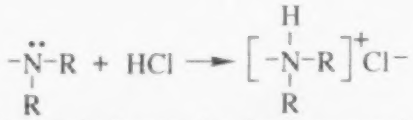
Sodium hydroxide is prepared by electrolysis of a concentrated aqueous sodium chloride solution.

$$2NaCl(aq) + 2H_2O(l) \rightarrow 2NaOH(aq) + Cl_2(g) + H_2(g)$$

TO THE CONTEMPORARY WORLD

1. **Drugs**

Lidocaine and novocaine are bases that are used by the medical profession as anesthetics. Because of their limited solubility in water (the typical injection solvent), the basic amino group in these molecules is converted to their hydrochloride salt form, which is water soluble. Furthermore, as bases these drugs would cause localized changes in pH at the injection site. Thus, the use of salt solutions minimizes this localized irritation.



$$-\underset{R}{\underset{|}{\ddot{N}}}-R + HCl \longrightarrow \left[-\underset{R}{\underset{|}{\overset{\overset{H}{|}}{N}}}-R\right]^{+}Cl^{-}$$

Antacids are basic compounds that decrease the amount of hydrochloric acid in the stomach. Examples are milk of magnesia (magnesium hydroxide), $Mg(OH)_2$, and Rolaids™ (sodium dihydroxy aluminum carbonate, $NaAl(OH)_2CO_3$). Sodium bicarbonate (baking soda), $NaHCO_3$, is the antacid in Alka Seltzer™. When Alka Seltzer™ is placed in water, bicarbonate ions react with hydronium ions from the acid producing the familiar fizz:

$$HCO_3^-(aq) + H_3O^+(aq) \rightarrow CO_2(g) + 2H_2O(l)$$

2. **Foods**

Acids and bases are used in cooking. Baking soda is sodium bicarbonate, $NaHCO_3$. Baking powder is a mixture of sodium bicarbonate, $NaHCO_3$, and an acid salt, potassium hydrogen tartrate, $KHC_4H_4O_6$. When bread bakes, these ingredients react to produce carbon dioxide and steam that expands and produces an airy loaf.

$$NaHCO_3(s) + KHC_4H_4O_6(s) \rightarrow H_2O(g) + CO_2(g) + NaKC_4H_4O_6(s)$$

Vinegar (acetic acid, $HC_2H_3O_2$) is a component of salad dressings. Citric acid is a component of fruit juices. Salts of acids (sodium benzoate and sodium propionate) are food additives that preserve foods from microorganisms.

3. **Agriculture**

The carbon in decaying organic matter uses oxygen to form carbon dioxide. This increased concentration of carbon dioxide causes the ground water flowing through to become acidic. Acidic soils are sour soils.

$$CO_2(g) + H_2O(l) \rightarrow H^+(aq) + HCO_3^-(aq)$$

Limestone (calcium carbonate, $CaCO_3$) is applied to acid soils to produce a pH that is slightly basic. Such a soil is called a sweet soil.

$$CO_3^{2-}(aq) + H^+(aq) \rightarrow HCO_3^-(aq)$$

4. **Insects**

Some species of millipedes store the compound mandelonitrile. When attacked by a predator, the millipede combines an enzyme with the mandelonitrile and discharges a toxic repellant, hydrogen cyanide, HCN.

$$C_6H_5CH(OH)CN(s) \rightarrow C_6H_5CHO(l) + HCN(g)$$

The whip scorpion, which is only a few centimeters long, can accurately direct a spray many times that distance at a predator. The spray contain a mixture of three organic carboxylic acids:

$HCOOH$	CH_3COOH	$CH_3(CH_2)_6COOH$
Formic acid	Acetic acid	Caprylic acid

Although acetic acid is the actual irritant, its effectiveness is enhanced by the presence of caprylic acid. Caprylic acid permits penetration of the waxy coating found on many recipients.

Formicidae ants defend themselves by discharging a spray of formic acid.

5. **Community Resources**

Try to plan field trips to a water or sewage treatment plant, a food processing plant, or dairy in your area. Invite knowledgeable people in your area associated with one of these to your classroom to discuss the chemistry involved in their particular plant.

6. **Acids and Foods**

Lactic acid, sauerkraut, hot BBQ sauce—all find their way to our tables. The relationship between acids and foods is a rich one to explore. (Idea contributed by Angie Matamoros.)

Extensions

Conjugate Acid-Base Pairs

The Brønsted-Lowry definition of an acid is a species that donates protons, and a base is a species that accepts protons. These definitions have a valuable application in interpreting experimental facts. The dissociation of an acid (or a base) is an equilibrium reaction. For example, in the dissociation of acetic acid,

$$HC_2H_3O_2(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$$

acetic acid donates a proton to water. Acetic acid is a Brønsted acid. Water, which accepts the proton, is a base.

In the reverse reaction (which is proceeding at the same rate), hydronium ion donates a proton to acetate ion. Hydronium ion is an acid. Acetate ion is a base because it accepts a proton.

Acetic acid by losing a proton forms a base, acetate ion. Acetate ion, by accepting a proton, can form an acid—acetic acid. This relationship can be represented as follows:

$$\underset{\text{Acid}}{HC_2H_3O_2} \quad \underset{\text{Gains } H^+}{\overset{\text{Loses } H^+}{\rightleftharpoons}} \quad \underset{\text{Base}}{C_2H_3O_2^-}$$

This acid-base pair, formed from each other by the loss or gain of a proton represents a conjugate acid-base pair. Conjugate means joined in pairs, or coupled.

Similarly, water (a base) accepts a proton to form the acid hydronium ion, which in turn, forms water by loss of a proton.

$$\underset{\text{Base}}{H_2O} \quad \underset{\text{Loses H+}}{\overset{\text{Gains } H^+}{\rightleftharpoons}} \quad \underset{\text{Acid}}{H_3O^+}$$

Hydronium ion and water are a second conjugate acid-base pair in the acetic acid dissociation.

Hydrolysis

When a salt dissolves in water, it completely dissociates to produce cations and anions. Either one or both of these ions may then react with water. This reaction is called hydrolysis. Hydrolysis of a salt can affect the pH of the solution. The resulting pH depends on the nature of the salt dissolved in water (see Figure 20).

Type of salt derived from	pH	Example
Strong acid-strong base	7	$NaCl$, KNO_3
Weak acid-strong base	>7	$NaC_2H_3O_2$, KCN
Strong acid-weak base	<7	NH_4Cl, $AlCl_3$
Weak acid-weak base	Depends on salt	$NH_4C_2H_3O_2$

Figure 20. pH effects by hydrolyzing typical salts.

Sodium acetate forms a basic solution because the acetate ion reacts with water to form hydroxide ion:

$$\underset{}{C_2H_3O_2^-} + H_2O \rightleftharpoons HC_2H_3O_2 + \underset{\text{Base}}{OH^-}$$

Similarly, ammonium chloride forms an acidic solution because only the cation hydrolyzes:

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + \underset{\text{Acid}}{H_3O^+}$$

Suggested Demonstration

Materials

Universal indicator

Five 0.1 M solutions selected from: sodium bicarbonate, $NaHCO_3$; ammonium chloride, NH_4Cl; sodium bisulfate, $NaHSO_4$; sodium chloride, $NaCl$; sodium carbonate, Na_2CO_3; ammonium sulfate, $(NH_4)_2SO_4$; aluminum nitrate, $Al(NO_3)_3$; sodium nitrate, $NaNO_3$; ammonium acetate, $NH_4C_2H_3O_2$; and potassium phosphate, K_3PO_4. Select solutions that will give acidic, basic, and at least one neutral solution.

5 Petri dishes or small beakers

Directions

Add solutions to vessels placed on the overhead projector. Add 1-3 drops of universal indicator solution to each container. Alternatively, large test-tubes in a test-tube rack may be used with a diffuse light source behind them.

Buffers

A buffer is a system that maintains a nearly constant pH in a solution when relatively small amounts of acid or base are added to the solution. A buffer is a mixture of a weak acid and its conjugate base, or a weak base and its conjugate acid.

A buffer can be prepared by mixing a weak acid (for example, carbonic acid, H_2CO_3) with one of its salts (for example, sodium bicarbonate, $NaHCO_3$), since the anion (HCO_3^-) of the acid is the conjugate base. This buffer is very important in maintaining a nearly constant pH of the blood. A weak base (NH_3) mixed with one of its salts (NH_4Cl) can also function as a buffer.

Suggested Demonstration

Materials

3 M or 6 M Hydrochloric acid, HCl (25 or 50 mL 12 M HCl diluted to 100 mL with distilled H_2O)

3 M or 6 M Sodium hydroxide, NaOH (12 g or 24 g NaOH(s) dissolved in 100 mL with distilled H_2O)

9 Large (about 180-mm) test-tubes

Test-tube rack with illuminated lightbox or white background

Universal indicator solution

Distilled water

Buffer Solution A: Prepare by mixing equal volumes of 4 M acetic acid, $HC_2H_3O_2$ (24 mL 17 M $HC_2H_3O_2$ diluted to 100 mL with distilled H_2O) and 4 M sodium acetate, $NaC_2H_3O_2$ (32.8 g $NaC_2H_3O_2$ dissolved in 100 mL distilled H_2O).

Buffer Solution B: Prepare by mixing equal volumes of 4 M aqueous ammonia, NH_3 (27 mL 15 M NH_3 diluted to 100 mL with distilled H_2O) and 4 M ammonium chloride, NH_4Cl (21.4 g NH_4Cl dissolved in 100 mL distilled H_2O).

Directions

Set up a 3 x 3 matrix of test-tubes.

HCl	Control	NaOH	
O	O	O	**Buffer A**
O	O	O	**Water**
O	O	O	**Buffer B**

Add Buffer Solution A (acetic acid buffer) to three test-tubes. Place distilled water into the second tube of each set. Then add Buffer Solution B (ammonia buffer) to the last three test-tubes. Add a drop or two of universal indicator solution to each tube; stir with a stirring rod until a brightly colored solution results. To the first of each of the three test-tube sets, add a dropperful of 6 M HCl; to the right hand member of each of the three test-tube sets, add a dropperful of 6 M NaOH. Stir the solutions. Have students record observed changes. To illustrate the capacity of the buffer, slowly add 6 M NaOH to the third test-tube in the acid buffer set (Set A) and 6 M HCl to the third test-tube in the basic buffer set (Set B) until pronounced color changes are noted.

Optional

Dissolve an Alka-Seltzer™ tablet in water in each of four 100-mL beakers on an overhead projector. Add universal indicator solution to each. Reserve the first beaker as a control. Add 1 M HCl dropwise to the second beaker with stirring; observe. Then add 1 M NaOH dropwise to the third beaker with stirring; observe. Finally, add distilled water dropwise to the fourth beaker with stirring; observe.

Lewis Acids and Bases

The Brønsted-Lowry concept of an acid and a base was limited to electron pair sharing to the proton (H^+). A new concept of acids and bases was proposed in the early 1920s by G. N. Lewis, a professor of chemistry at the University of California at Berkeley. In the Lewis proposal, an acid is an electron-pair acceptor ("take two from you"), a base is an electron-pair donor ("have pair, will share"), and an acid-base reaction involves a base sharing an electron pair with an acid.

Lewis extended the Brønsted-Lowry definition of an acid and a base. Lewis acids and bases are not dependent upon the proton (H^+) or the hydroxide ion. In fact, the Lewis concept revolutionized the theory and practice of acid catalysis in organic chemistry (see last equation in Figure 21, formation of a carbocation).

In a Lewis acid-base reaction a coordinate covalent bond is formed between an acid and base. Examples are in the equations shown in Figure 21:

Acid + Base

$$F_3B + :NH_3 \rightarrow F_3B:NH_3$$

$$Ag^+ + 2:NH_3 \rightarrow [H_3N:Ag:NH_3]^+$$

$$AlCl_3 + (CH_3)_3C\text{-}Cl \rightarrow [Cl:AlCl_3]^- + (CH_3)_3C^+$$

Figure 21. Examples of Lewis acid-base reactions.

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A BASIC program for use on the IBM-PC that will generate on screen any desired acid-base master variable diagram and will generate a hard copy on request.

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This publication by the American Chemical Society describes fun physical science activities on acids and bases for children and adults to do together.

Bell, R. P. (1973). *The proton in chemistry* (2nd Ed.). New York, NY: Cornell University Press.

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Borgford, C., and Summerlin, L. (1988). *Chemical activities*. Washington, DC: American Chemical Society.

A variety of activities, with teaching tips and safety discussions, useful as demonstrations or laboratory experiments for grades 7-12.

Brooks, D. W. (Producer). (1989). *Doing chemistry* [videodiscs, computer program and supporting written materials]. Washington, DC: American Chemical Society.

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Gadek, F. (1987). A commercially available device for conductivity experiments. *Journal of Chemical Education, 64,* 281-282.

This conductivity tester uses the continuity and tone-generating chassis from Radio Shack. The chassis uses a 9-V transistor battery.

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This device can be made from a discarded 35-mm film canister, 9-V battery replacement snaps, a 200-300 ohm resistor, and a light emitting diode.

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A versatile, comprehensive set of programs useful to students and teachers in any course that deals quantitatively with acid-base equilibrium. The "Acid-Base Package" runs under MS-DOS or IBM-PC DOS Version 3.0 or higher. It is supplied on 5.25-in disks and requires at least 512 K RAM.

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A collection of activities from the Institute for Chemical Education workshops readily usable by K-12 teachers.

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Many demonstrations with procedures and detailed explanations.

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Summerlin, L., and Ealy, J. (1985). *Chemical demonstrations: A sourcebook for teachers, Volume 1.* Washington, DC: American Chemical Society.

Two good compilations of demonstrations for teachers of chemistry and physics.

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This entire issue contains articles dealing with swimming pool chemistry, antacids, acid-base indicators (a laboratory activity), pH and hair shampoo, acid rain, acids and bases: ancient concepts in modern science.

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Describes construction of a conductivity apparatus that allows quantification of conductivity and avoids using a 120-V light bulb and bare electrodes.

Appendix

- **Transparency Masters**
  1. Acids: Concentration *vs.* Strength
  2. Acid-Base Definitions
  3. pH and Concentration
  4. Word Search
- **Humor**

Acids: Concentration *vs.* Strength

	CONCENTRATED	DILUTE
STRONG	$H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ A^- **HA** $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ **HA** $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ **HA** H^+ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ **HA** $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ **HA** $H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$ $H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-$ $A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+\ A^-\ H^+$	$H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+$ $A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-$ $H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+$ $A^-\quad H^+$ HA $A^-\quad H^+\quad A^-$ $H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+$ $A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-$ $H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+$ $A^-\quad H^+\quad A^-\quad H^+\quad A^-\quad H^+\quad A^-$
WEAK	HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA **$H^+\ A^-$** HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA **$H^+\ A^-$** HA HA HA HA HA **$H^+\ A^-$** HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA HA HA HA HA **$H^+\ A^-$**	HA HA **$H^+\ A^-$** HA HA HA HA **$H^+\ A^-$** HA HA HA HA HA HA HA HA HA HA HA HA **$H^+\ A^-$** HA HA HA HA **$H^+\ A^-$** HA HA

Acid-Base Definitions

	Arrhenius	**Brønsted-Lowry**	**Lewis**
Acid	Provider of H^+ in HOH	Proton donor	Electron pair acceptor
Base	Provider of OH^- in HOH	Proton acceptor	Electron pair donor
Neutralization	Formation of water	Proton transfer	Coordinate covalent bond formation
Equation	$H^+ + OH^- \rightarrow HOH$	$HA + B \rightarrow BH^+ + A^-$	$A + :B \rightarrow A{:}B$
Limitation	Water solution only	Proton transfer Reactions only	Generalized theory

pH and Concentration (25 °C)

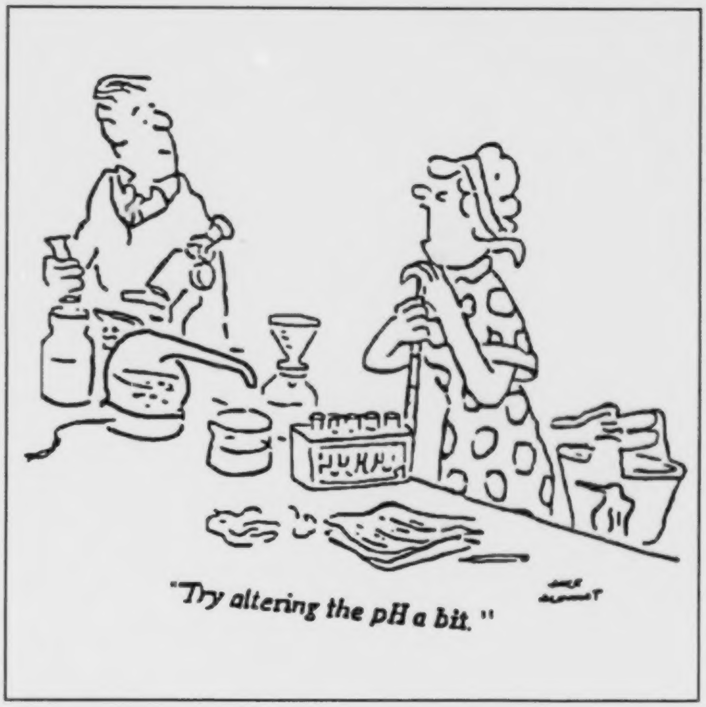
	$[H^+]$	pH	$[OH^-]$	pOH
Acidic Solution	$>10^{-7}$	<7	$<10^{-7}$	>7
Neutral Solution	10^{-7}	7	10^{-7}	7
Basic Solution	$<10^{-7}$	>7	$>10^{-7}$	<7

Word Search

Z S N M A T K B Z R U O S T M P Y
X T S Q R A D H V K Q K Y O S E W
R P H T D R B U X E R I L D U R H
E U Y O R U M X J O S A I Q H B Y
T R D F H E P G N H R W N A V W D
T M R Z D V N O Q I O P O Y K M R
I R O C V M D G T R T T I X D O O
B I G V F N J Y T P A V W L X D X
A U E A O X Y N W H C C R Z J S I
M Y N T K A T H T N I O P D N E D
G Y O P V T E I D Z D F E S I K E
M R O Y D I A V C M N G J T H S J
P Q Z N E U T R A L I Z E I N I R

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

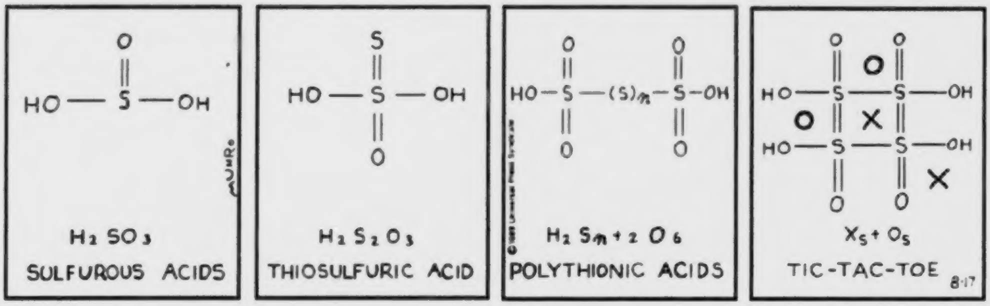
1. Weak acids or bases that change color over a narrow pH range.
2. To eliminate the acidic or basic character of a solution.
3. In a solution, the percentage of ionized molecules of acid or base.
4. The point in a titration when the indicator changes color (2 words).
5. Concentration of a solution defined in terms of moles of solute per liter of solution.
6. Ion most often associated with an acid.
7. Ion most often associated with a base.
8. Taste most often associated with acids.
9. Taste most often associated with bases.
10. Brønsted-Lowry acid (2 words).


"Try altering the pH a bit."

CHEMTECH, March 1975, *5*(3), 164. © American Chemical Society. Reprinted with permission.

EUREKA!

by Munro Ferguson


HO — S — OH
H_2SO_3
SULFUROUS ACIDS
HO — S — OH
$H_2S_2O_3$
THIOSULFURIC ACID
HO—S—(S)n—S—OH
$H_2S_{n+2}O_6$
POLYTHIONIC ACIDS
HO—S—S—OH
HO—S—S—OH
$X_5 + O_5$
TIC-TAC-TOE
8-17

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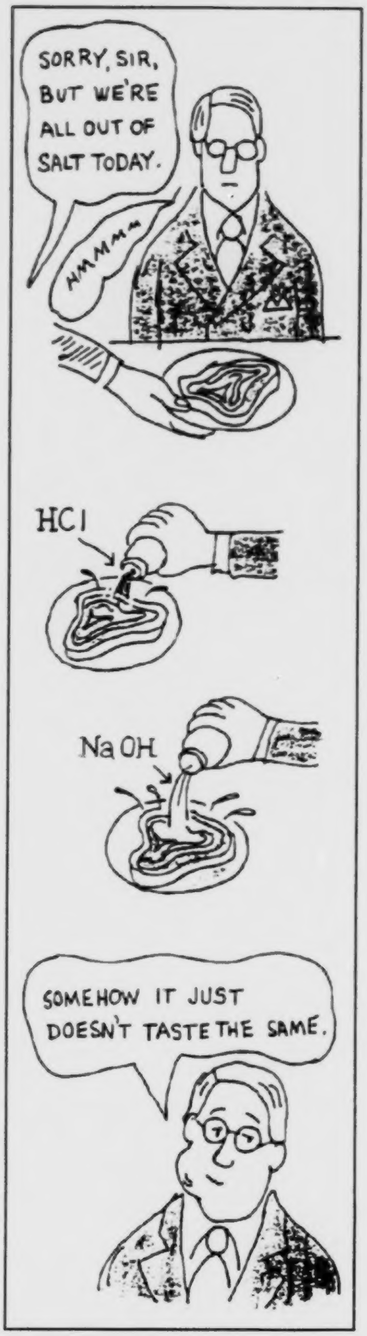
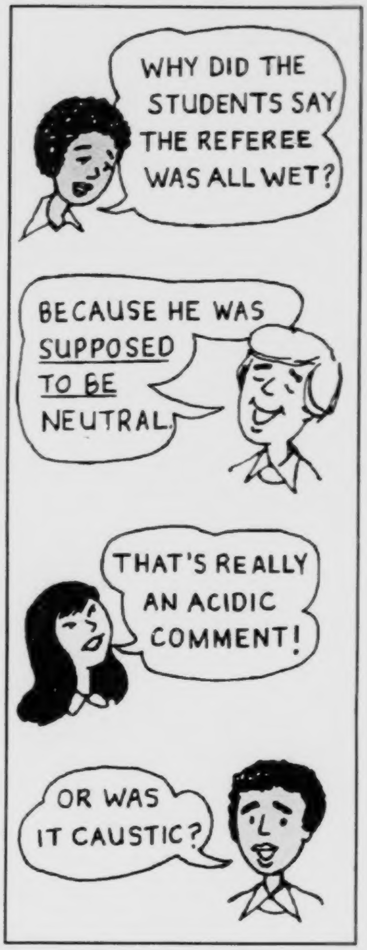

SORRY, SIR, BUT WE'RE ALL OUT OF SALT TODAY.
HMMMM
HCl
NaOH
SOMEHOW IT JUST DOESN'T TASTE THE SAME.

Used by permission from Veritex Publishing Co.


WHY DID THE STUDENTS SAY THE REFEREE WAS ALL WET?
BECAUSE HE WAS SUPPOSED TO BE NEUTRAL.
THAT'S REALLY AN ACIDIC COMMENT!
OR WAS IT CAUSTIC?

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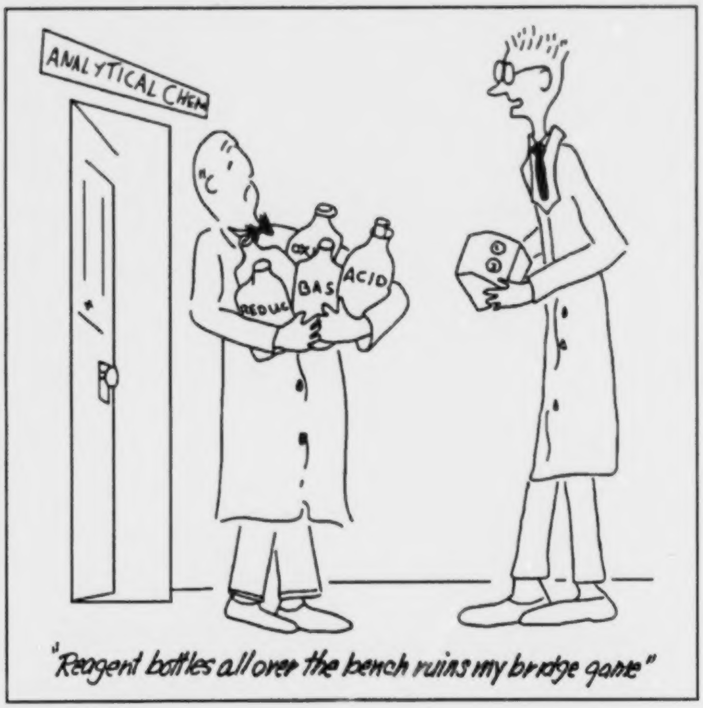
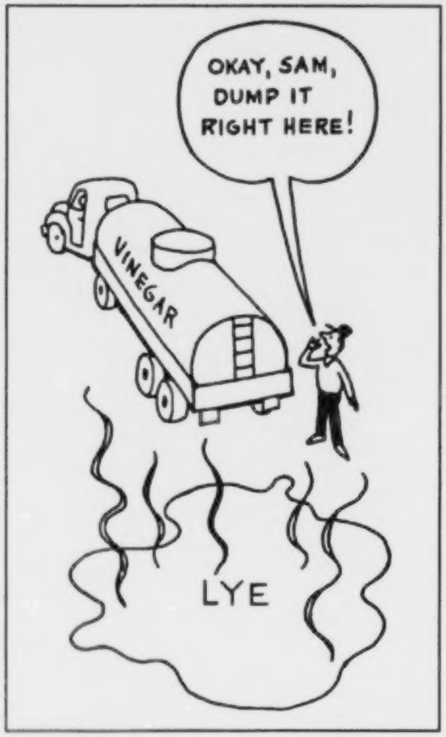

ANALYTICAL CHEM
REDUC
BAS
ACID

"Reagent bottles all over the bench ruins my bridge game"

CHEMTECH, December 1972, p. 737. © American Chemical Society. Reprinted with permission.


OKAY, SAM, DUMP IT RIGHT HERE!
VINEGAR
LYE

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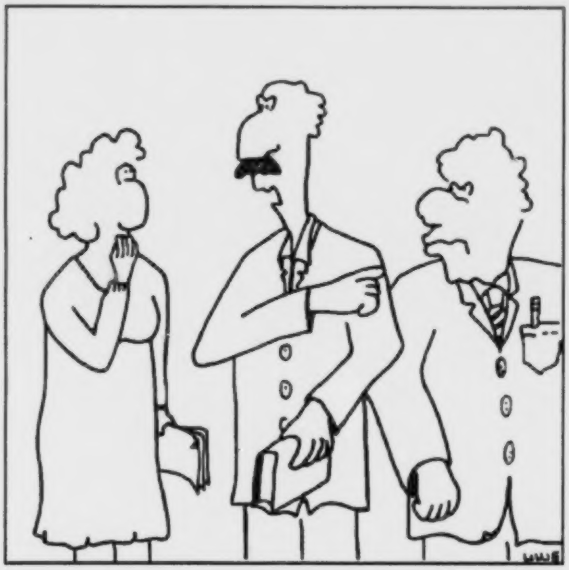
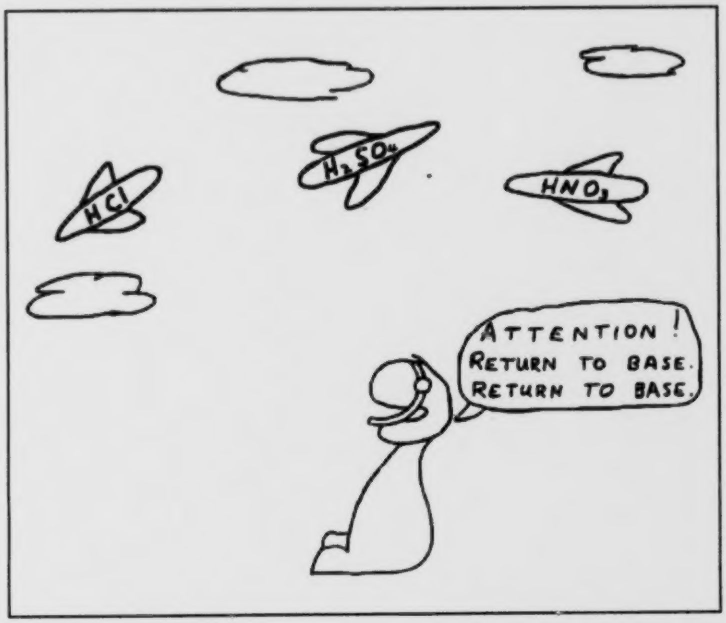
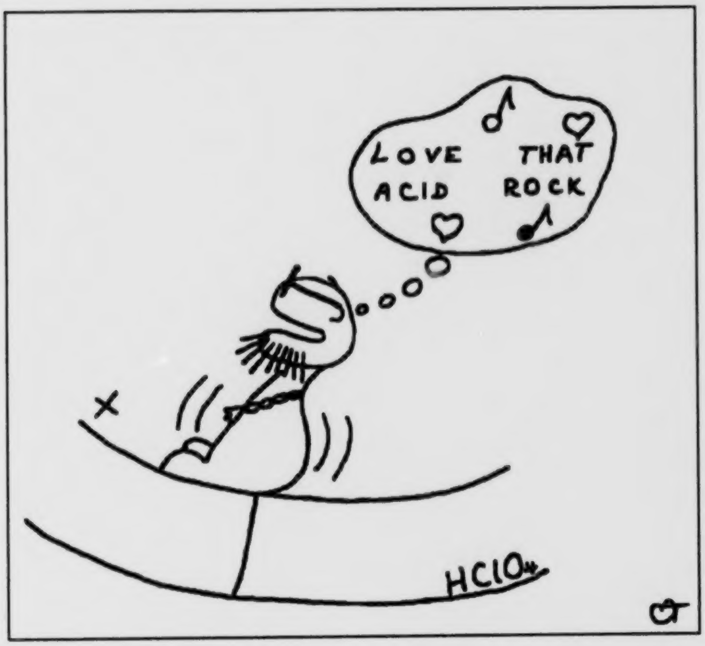


WITH HIS ACID TONGUE AND MY CAUSTIC WIT, WORKING IN THE SAME LAB COULD END IN ASSAULT.

CHEM 13 NEWS, April 1980, p. 3. Reprinted with permission.


HCl
H_2SO_4
HNO_3
ATTENTION! RETURN TO BASE. RETURN TO BASE.

CHEM 13 NEWS, November 1983, p. 9. Reprinted with permission.


LOVE THAT ACID ROCK
$HClO_4$

CHEM 13 NEWS, December 1983, p. 16. Reprinted with permission.

A SourceBook Module

Version 2.0 1994

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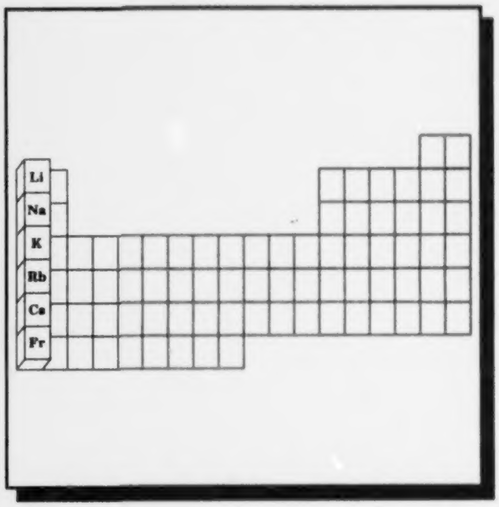


ALKALI METALS

ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

From saline solutions to street lamps, from baking powders to bleaches, from fertilizers to fireworks, alkali metal compounds find everyday uses in our lives. Elements necessary for life, yet ones that react violently with water: do these two sound incompatible? In their aqueous ionic forms, sodium and potassium ions (Na^+ and K^+) are essential ingredients for animal and plant life, yet the elemental atomic forms of these alkali metals react vigorously with water and other compounds.

The alkali metals (Li, Na, K, Rb, Cs, Fr) form a vertical family of elements that begin each horizontal row of the Periodic Table. Francium, element 87, is a rare radioactive decay product of the radioactive element, actinium. Even the most stable isotope of francium decays so rapidly that its chemical properties are not well known. The other alkali metals are silvery metallic solids, as soft as cold butter. Lithium, sodium, and potassium are less dense than water; hence they float on it (but don't try it!). Lithium's density (0.53 g/cm^3) is just over half that of water. All have low melting points. Cesium melts at 29 °C, a bit above room temperature; Na melts at 98 °C, just below the boiling point of water. The alkali metals are so highly reactive that they never occur free (in their elemental form) in nature, but always in combination with other elements. Since they react rapidly with oxygen in air and violently with water, they must be stored in unreactive oil or kerosene.

All alkali metals react vigorously with halogens to produce alkali halides. Large quantities of alkali metal chlorides are found in the oceans, inland seas, and salt deposits. Fifty million billion (5×10^{16}) tons of salt (NaCl) are dissolved in earth's oceans. The sodium ion, Na^+, is the principal positive ion in fluids surrounding cells in our bodies, where it is needed for water retention and muscle action. This ion (in the form of saline solution) is often given intravenously to hospital patients. Potassium ions, K^+, are also essential to life, both plant and animal. Compounds of this ion, such as KCl, K_2SO_4, and KNO_3, are used extensively as fertilizers (see *Industrial Inorganic Chemistry* module). Insoluble lithium carbonate, Li_2CO_3, is used to treat manic depressives, although its mode of action is not well understood.

Each alkali metal atom has one more electron than the chemically stable noble gas atom just preceding it in the Periodic Table. Each atom has a large relative size (radius), coupled with filled inner energy levels of electrons. Each atom can therefore readily lose this one electron, forming stable +1 ions with noble gas electron configurations. Thus they have low ionization energies. Cesium's ionization energy is so low that visible light can ionize it, permitting its use in photoelectric cells, where light energy is converted directly into electricity.

The chemistry of alkali metals provides a fascinating entry into the field of descriptive chemistry and a perfect introduction to the concept of periodicity of the elements (see *Periodicity* module).

PLACE IN THE CURRICULUM

The similarities in physical and chemical properties of alkali metals reinforce the concept of families of elements and thus serve as an excellent introduction to the Periodic Table (*Periodicity* module.) Due to their strong tendency to form compounds in which they exhibit only the +1 oxidation state, their chemistry is simple and predictable, yet frequently exciting. The topic could thus serve as an introduction to predicting the products of chemical reactions (*Simple Chemical Reactions* module).

As a consequence of their high solubility in water and their relative abundance in nature, the compounds of sodium and potassium find extensive use. Examples familiar to students include table salt (NaCl), baking soda ($NaHCO_3$), lye (NaOH), potash (K_2CO_3), soap ($C_{17}H_{35}COONa$), and detergents (*e.g.*, $C_{12}H_{25}OSO_3^-Na^+$). All these reasons suggest introducing the topic of alkali metals early. The extraction of the alkali metals from their naturally occurring compounds provides a simple introduction to the concept of oxidation-reduction reactions (*Oxidation-Reduction* module), while much of the chemistry of the anions in the alkali metal compounds focuses on acid-base (*Acid-Base* module), precipitation (*Solubility and Precipitation* module), and redox reactions. Alkali metals can thus be related to several major areas of the typical high school chemistry curriculum.

CENTRAL CONCEPTS

1. The alkali metal family, consisting of Li, Na, K, Rb, Cs, and Fr, is a highly reactive family of elements. The elements in this family are never found free in nature (as M^o atoms), existing instead in chemical combination with anions as ionic compounds (containing M^+ ions).
2. The alkali metals are prepared by electrolytic (electrolysis) or chemical reduction [$M^+ + e^- \rightarrow M^o$].
3. Alkali metal cations have a low charge density and relatively weak attractions for negative ions (anions). As a consequence, most salts of alkali metal ions are very water-soluble, and their hydroxides are very soluble strong bases (alkalis).
4. The attractive forces among atoms in alkali metal crystals are relatively weak. As a result each metal has low density, melting point, heat of fusion, electronegativity, ionization energy, and electron affinity; they are also soft, malleable, and ductile.
5. The chemical and physical properties of these elements are similar (as are those of their ions), and most of their properties change in a regular periodic manner as one goes down the family in the Periodic Table.
6. All reactions of the alkali metal *elements* (as well as those of almost all elements) are oxidation-reduction reactions.
7. The vast majority of reactions of alkali metal *compounds* are acid-base reactions involving the anions associated with the alkali metals. In oxidation-reduction and precipitation reactions involving alkali metal compounds, it is almost always the anion originally associated with the alkali metal that is involved in the observed net reaction; *e.g.*,

$$AgNO_3(aq) + NaCl(aq) \rightarrow NaNO_3(aq) + AgCl(s)$$

8. All alkali metals form body-centered cubic crystals.
9. All alkali metals react vigorously with water, producing the alkali metal hydroxide and hydrogen gas.

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$

10. Alkali metals' relatively large atomic size (radii) and single electron far from the nucleus, account for their low electron densities and hence low ionization energies. This observation can be explained using Coulomb's Law [$F = \frac{q^+ q^-}{kr^2}$], where the force of attration (F) is directly proportional to charge (q) and inversely proportional to atomic radius (r).

RELATED CONCEPTS

1. Chemical symbolism and nomenclature, formulas, atomic and molar masses
2. Equation writing, types of reactions, predicting reaction products
3. Stoichiometry
4. Solubility equilibrium
5. Solubility
6. Chemical periodicity
7. Atomic structure, electronic structure, Lewis structures
8. Molecular structure, ionic compounds
9. Crystal structure (ionic, metallic)
10. Acid-base reactions
11. Precipitation reactions
12. Oxidation-reduction reactions, electrolysis

RELATED SKILLS

1. Manipulative skills:
    a. The ability to write chemical formulae, predict reaction products, and balance chemical equations.
    b. An understanding of basic techniques involved in qualitative analysis of metal ions.
    c. An awareness of special precautions involved in lecture demonstrations and laboratory activities involving free alkali metals and their corrosive compounds.
2. Mathematical and conceptual skills:
    a. Mastery of stoichiometric skills, unit analysis, solution concentrations.
    b. Mastery of arithmetic skills, including simple algebra, use of exponents, and roots.
    c. Working knowledge of Coulomb's law and parameters involved in electrical attractions.
    d. Ability to write/draw electron configurations.
    e. Ability to draw Lewis-dot structures.
    f. Ability to predict bond and molecular polarity on the basis of ionization energy (or electronegativity) and molecular geometry.
    g. Ability to predict trends in properties of alkali metals (chemical periodicity) on the basis of atomic structure.

PERFORMANCE OBJECTIVES

After completing their study of alkali metals, students should be able to:

1. explain why alkali metals are never found in pure metallic state in nature.
2. describe the primary ways in which alkali metals are prepared or obtained.
3. explain why most alkali metal compounds are water-soluble.
4. discuss physical properties of alkali metals in terms of atomic structure.

5. describe trends in physical and chemical properties as related to alkali metals.
6. explain why many reactions involving elemental alkali metals are redox reactions, but reactions of compounds of these elements rarely involve changes in oxidation state of alkali metal ions.
7. explain why reactions involving compounds of alkali metals are almost all acid-base reactions of anionic constituents.
8. describe crystal structure of alkali metals.
9. explain physical properties of the alkali metals in terms of their crystal structures.
10. explain the relatively low ionization energies of alkali metals as they relate to Coulomb's Law.

Concept/Skills Development

LABORATORY ACTIVITIES

Chemical properties of free alkali metals do not lend themselves to hands-on activities for high school laboratories due to the hazardous nature of these elements. With great care, you can carry out lecture demonstrations with selected alkali metals. Several demonstrations are detailed in the next section. Activities involving alkali metal compounds do exist and are available in many laboratory manuals. Some chemistry of these metals is also covered in the demonstration section of this module.

Alternatively, audio-visual aids such as the CHEM Study film, "Chemical Families"; the Project Seraphim Periodic Table Videodisc; available video lecture demonstration tapes published as companions to new college general chemistry texts; or the "Doing Chemistry" videodisc program can serve as excellent substitutes for the hazardous live activities.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Identification of Metal Ions by Flame Tests

Materials

Some of the following ionic solids:
- Barium chloride, barium hydroxide*, barium nitrate
- Calcium carbonate, calcium chloride*, calcium nitrate*
- Cesium carbonate*, cesium chloride*, cesium nitrate
- Copper(II) acetate*, copper(II) chloride*, copper(II) sulfate
- Lithium carbonate, lithium chloride*, lithium nitrate
- Potassium nitrate, potassium carbonate, potassium chloride*
- Rubidium carbonate*, rubidium chloride, rubidium hydroxide*
- Sodium chloride, sodium nitrate*
- Strontium chloride*, strontium nitrate*
- *Somewhat soluble in 95% ethanol (see *Procedure B*)

Striker or matches
Burner
Safety goggles
Small glass bottles with screw caps for ionic solids above
Pump hair spray bottles
Petri dishes
Ethanol or methanol

Safety

Try all demonstrations ahead of time to ascertain the safest reaction conditions. Be aware that any of these methods allows metal compounds to be released into the air. Use adequate ventilation. Always wear safety goggles for chemical demonstrations.

Directions (Choose either Procedure A, B, or C)

Procedure A: Several Powdered Solids

Select several metallic salts from the preceding list. Powder each by crushing a one-tablespoon sample in a clean mortar and pestle. Place each powder in a separate labeled, tightly capped bottle (100-mL or larger). Set up a burner, making sure the flame is the normal light blue. Vigorously shake each bottle one at a time to create a finely divided solid mist, and remove the cap while holding the mouth of the bottle upright and close to the burner's air intake. A prolonged colored flame will be observed. If the procedure is carried out in a darkened room and students are provided with inexpensive plastic diffraction gratings (available as 2" x 2" mounted slides from Edmund Scientific, Great Barrington, NJ 08007), they may be able to observe the line spectra of the metals.

Procedure B: Ethanol (or Methanol) Solutions

The solids in the preceding list that are marked with asterisks (*) are somewhat soluble in ethanol (or methanol) and are good choices for use in this procedure. Place a pea-sized quantity of several of these powdered solids in glass Petri dishes or small beakers. Cover the solid with a thin layer of ethanol or methanol. Darken the room if possible and ignite the alcohol with a long match or taper. The colors will continue as long as there is alcohol to burn.

Procedure C: Aqueous Solutions

Use pump spray bottles, obtainable in garden shops. Fill each bottle with dilute (0.1 to 0.5 M) solutions of selected compounds listed above. (You probably have some of these solutions already prepared on your shelves.) Spray the liquids into the burner flame, one at a time. Observe instant, bright bursts of color.

Discussion

The colors of the flames will be as follows:

Barium compounds	(Ba^{2+})	Apple green
Calcium compounds	(Ca^{2+})	Orange-red
Cesium compounds	(Cs^{+})	Blue
Copper compounds	(Cu^{2+})	Green
Lithium compounds	(Li^{+})	Red
Potassium compounds	(K^{+})	Violet (lilac)
Rubidium compounds	(Rb^{+})	Purple
Sodium compounds	(Na^{+})	Yellow
Strontium compounds	(Sr^{2+})	Deep red

An alternative procedure that works well is the use of carbonate or oxide compounds of calcium, copper, lithium, sodium, and potassium in empty film canisters. Shake the closed canister, then open the cap near the base of the burner. The dust from the canister will enter the fuel flow and turn the flame a brilliant color.

Demonstration 2: Reactions of Metals with Water

You should perform these activities as a demonstration rather than as a student activity because of the high reaction rates, the somewhat variable way in which the reactions occur, and the dangers involved in handling sodium. Because of the danger of storing and handling potassium, it is recommended that a video of potassium reactivity be used along with this demonstration to extend the periodic trend in reactivity.

Purpose

To demonstrate activity level of alkali and alkaline earth metals in the presence of water.

Materials

Lithium metal, small piece
Sodium metal*, small piece
Potassium metal, small piece
Calcium metal, small piece
Magnesium ribbon, 6-10 cm
Water

Liquid detergent (not basic), few drops
Phenolphthalein solution (0.01 g phenolphthalein per 100 mL ethanol)
Beaker
4 Glass Petri dishes or crystallizing dishes
Fine mesh wire gauze
Test-tube and stopper
Funnel (glass)
Overhead projector
*Dri-Na works well here in place of sodium metal. Dri-Na is sold by Flinn Chemical Company. It is an alloy containing 12% sodium with lead. This alloy generates hydrogen gas slowly and safely.

Safety

Sodium and potassium are potentially explosive if large chunks are used, or if a peroxide coating exists on the surface of the metal. Use only freshly cut pieces of shiny metal. You might consider using Dri-Na in place of the sodium metal.

Procedure

Part 1

1. Fill a glass vessel (Petri dish or beaker) half-full with water. Add three drops of phenolphthalein and three drops of neutral detergent. (Place the Petri dish on overhead projector.)
2. Drop in a small piece of lithium and place the wire gauze over the top of the glass container (or if using the overhead, use the cover of the Petri dish).
3. Repeat Steps 1-2 with sodium, potassium and calcium. Note that calcium oxidizes easily and reacts vigorously with water. Purchase only small quantities of calcium and keep tightly covered when not in use. If your calcium is white and powdery, it has oxidized. Choose only metallic pieces of calcium for the above test.

Discussion

The equations for the reaction are:

$$\begin{aligned}
2\mathrm{Li}(s) + 2\,\mathrm{H_2O}(l) &\rightarrow 2\mathrm{LiOH}(aq) + \mathrm{H_2}(g)\\
2\mathrm{Na}(s) + 2\,\mathrm{H_2O}(l) &\rightarrow 2\mathrm{NaOH}(aq) + \mathrm{H_2}(g)\\
2\mathrm{K}(s) + 2\,\mathrm{H_2O}(l) &\rightarrow 2\mathrm{KOH}(aq) + \mathrm{H_2}(g)\\
\mathrm{Ca}(s) + 2\,\mathrm{H_2O}(l) &\rightarrow \mathrm{Ca(OH)_2}(aq) + \mathrm{H_2}(g)
\end{aligned}$$

The trail of zigzagging sodium (potassium, calcium) should leave a pink trail as the indicator reacts with the metal hydroxide formed.

Part 2

1. Fill a test-tube completely with water. Place a 6-10 cm coil of magnesium ribbon into the test-tube mouth. Stopper the test-tube, invert it into the tall narrow beaker, which has been half-filled with water, and remove the stopper. Heat the water. Have students observe and record results.

or

2. Place a 6-10 cm coil of clean magnesium ribbon in a beaker half-filled with water. Invert a short stem glass funnel over the magnesium and place a test-tube over the funnel stem. Have students observe and record results. Next, heat the water to near boiling and have students record observations.

Discussion

Once again, magnesium readily forms oxide coating on its surface. The pieces of magnesium used for the demonstration should be cleaned beforehand by dipping all pieces in 1-3 M HCl for a short time and then rinsing with tap water and drying.

The equation for the reaction is:

$$Mg(s) + 2H_2O(l) \xrightarrow{heat} Mg(OH)_2(s) + H_2(g)$$

NOTE: This reaction happens very slowly, if at all, at room temperature, but bubbles of hydrogen are observed at elevated temperatures.

Compare and contrast the reactivity of the four metals with water. Have students explain the differences in terms of chemical periodicity.

[The revised *ChemStudy* film "Chemical Families," now available on videotape, beautifully demonstrates the reactivity of alkali metals with water and halogens. Many college texts now make available to teachers videotaped demonstrations that include reactions of alkali metals.]

Demonstration 3: Solvay Process Demonstration for the Commercial Preparation of $NaHCO_3$ and Na_2CO_3

Materials

Acetone, 5 mL
15 M (concentrated) ammonia, 70 mL
Sodium chloride, NaCl, 20 g
Dry ice, 100 mL (or marble chips and 6 M HCl to generate CO_2)
Water
Ice
Calcium hydroxide, $Ca(OH)_2$, saturated solution
Beaker, 100-mL, 400-mL, 800-mL
Buchner funnel, filter paper
Suction flask
3 Glass tubing with angle-bends
2 Gas-generating bottles
Thistle tube
Thermometer
Graduated cylinder, 50-mL
Burner or hotplate
Stirring rod
Erlenmeyer flask, 250-mL
Rubber tubing, 2 pieces
2 Stoppers to fit gas bottles, 2-hole
Hydrion paper, Universal indicator or pH meter

Safety

Always try demonstrations before doing them for a class. Wear safety goggles when doing any chemical demonstration. Concentrated ammonia is an eye, nose, and throat irritant. Be sure you have adequate ventilation. Dry ice is extremely cold and can cause burns. Handle with care. Use gloves.

Directions

1. Prepare a warm water bath by heating ~600 mL water in an 800-mL beaker to 50-60 °C.

2. Add 25 mL distilled water to 100-mL beaker and place in ice to serve as an ice water rinse. Reserve.
3. To 20 g sodium chloride (NaCl) in a 400-mL beaker in a hood add 5 mL portions of 15 M NH_3 with continuous stirring, until the NaCl has dissolved and a total of 65-70 mL of aqueous ammonia has been added.
4. The addition of carbon dioxide may be carried out by adding powdered dry ice or making a carbon dioxide gas generator using marble chips and 6 M HCl.

Dry Ice Method

5. Add 65-70 mL powdered dry ice to the NaCl-NH_3 mixture with continuous stirring until precipitation occurs. Warm the reaction vessel to 10-15 °C occasionally by immersing the bottom of the beaker in the warm water bath. Add an additional 20-25 mL of dry ice continuing to stir until the bubbling of carbon dioxide ceases. Proceed to Step 6.

Carbon Dioxide Generator Method

5. The apparatus is set up as shown. Use glycerol to lubricate the glass tubing and thistle tube before inserting into rubber stoppers. Always use a towel wrapped around the tubing to protect your hands. Place 60 g marble chips into the gas generating bottle. Place 25-35 mL of water into the bottle clamped to the ring stand so that the end of the glass tubing is below the water level. Add 6 M HCl and bubble the carbon dioxide into a 250-mL or 500-mL Erlenmeyer flask containing your reaction mixture. Swirl the flask to assist the mixing operation until a large quantity of solid forms. This method takes quite a bit of time (about 1.5 h). Complete the synthesis as indicated in Step 6.

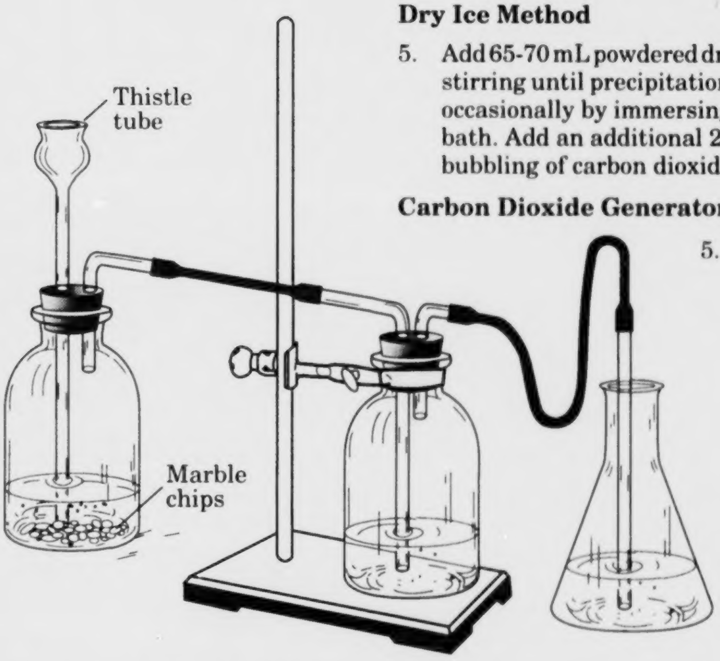


Figure 1. Carbon dioxide generating apparatus.

6. Cool the mixture in the ice bath. Suction filter the precipitate with a Buchner funnel fitted with filter paper and attached to a suction flask. With the suction off, pour 5 mL ice water over the precipitate, letting it soak into the precipitate. Reapply suction to remove the water rinse. Repeat the rinsing process with another portion of ice water. Repeat with 5 mL acetone. Pull air through the precipitate 5-10 min to initiate drying of the product. If desired, the product can be dried completely by transferring to an evaporating dish and drying at 110 °C for about two hours. (This is not necessary if you simply want to show the class the precipitate of $NaHCO_3$. If a quantitative determination of yield is to be made, it is helpful to know that dry ice has a density of about 1.5 g/mL.)

Testing the Product

7. Dissolve a pea-sized portion of the product in 5-10 mL distilled water. Determine the pH with hydrion paper, universal indicator or pH meter.
8. Place a portion of the solid in a test-tube and add 2-3 mL 6 M HCl. Test the evolving gas with an eyedropper containing a hanging drop of saturated calcium hydroxide solution. (The drop becomes cloudy due to a white precipitate of $CaCO_3$.)

Discussion

The equation for preparation of $NaHCO_3(s)$ is:

$$H_2O(l) + NH_3(aq) + CO_2(g) + NaCl(aq) \rightarrow NaHCO_3(s) + NH_4Cl(aq)$$

The solid sodium hydrogen carbonate formed is separated by decantation or filtration. In the Solvay Process, the compound is then dried and heated to 175 °C to produce Na_2CO_3.

$$2NaHCO_3(s) \rightarrow Na_2CO_3(s) + H_2O(g)$$

Reactions for *Testing the Product* are:

$$HCO_3^-(aq) + HOH(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + CO_2(g) + H_2O(l)$$

$$CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$$

Demonstration 4: Qualitative Analysis of Alkali Metal Ions

Materials

6 M Acetic acid, CH_3COOH, 10 mL [3.5 mL glacial (17 M) acetic acid diluted to 10 mL]

95% Ethanol, 20 mL (95 mL ethanol per 100 mL solution)

0.25 M Potassium nitrate, KNO_3 (2.5 g KNO_3 per 100 mL solution)

0.06 M Silver nitrate, $AgNO_3$ (1.7 g $AgNO_3$ in 100 mL solution)

Sodium bitartrate, $NaHC_4H_4O_6$, saturated solution, 2 mL (Add solid $NaHC_4H_4O_6$ to 2 mL H_2O with stirring and until no more dissolves.)

Sodium cobaltinitrite, $Na_3Co(NO_2)_6$, solution (Prepare in hood. Dissolve 0.75 g $Co(NO_3)_2{\cdot}6H_2O$ in 3.0 mL of water; dissolve 6.0 g sodium nitrite ($NaNO_2$) in 3.0 mL of water; mix the two solutions with vigorous stirring. Add 1.5 mL glacial acetic acid. Dilute to 25 mL, let stand, filter.)

0.1 M Sodium nitrate, $NaNO_3$ (3.2 g $NaNO_3$ per 100 mL solution)

Sodium perchlorate, $NaClO_4$, saturated solution, 2 mL (Add solid $NaClO_4$ to 2 mL H_2O with stirring and until no more dissolves.)

Zinc uranyl acetate, $Zn(UO_2)_3(C_2H_3O_2)_8$, solution (Dissolve 5 g uranyl acetate, $UO_2(C_2H_3O_2)_2{\cdot}2H_2O$, in 1 mL glacial acetic acid diluting to 25 mL with water. In a separate container stir 15 g zinc acetate with 1 mL of glacial acetic acid diluting to 25 mL with water. Mix the two solutions, add 0.5 g sodium chloride, let stand overnight, and filter.)

Ice bath

Safety

Sodium perchlorate is a highly reactive oxidizing agent. It is stable in aqueous solution and is traditionally used as a precipitant for K^+ in qualitative analysis. Chlorates ($MClO_3$) and perchlorates ($MClO_4$) in solid form should be kept away from organic and other combustible solids and should never be handled with metal spatulas.

Directions

Test for Potassium Ion

1. Place 3 mL $KNO_3(aq)$ in each of three test-tubes. To the first test-tube add the following: 3 mL 95% (by volume) ethanol; 1 mL 6 M acetic acid; 1 mL $AgNO_3$ solution; and 3 mL $Na_3Co(NO_2)_6$. Stir and cool in an ice bath. A yellow precipitate of $K_2AgCo(NO_2)_6$ indicates the presence of potassium ion. (The reaction would also precipitate NH_4^+, Li^+, and Tl^+.)
2. To the second test-tube add 3 mL 95% ethanol and 2 mL saturated $NaHC_4H_4O_6$. Stir and cool in an ice bath. A white precipitate of $KHC_4H_4O_6$ should form.

3. To the third test-tube add 3 mL 95% ethanol and 2 mL saturated $NaClO_4$. *Do not heat!* A white precipitate of $KClO_4$ will form upon cooling.

Test for Sodium Ion

Place 2 mL $NaNO_3(aq)$ in a test-tube. Add 6 mL zinc uranyl acetate, $Zn(UO_2)_3(C_2H_3O_2)_8(aq)$. Stir vigorously and cool in an ice bath. A greenish-yellow precipitate of $NaZn(UO_2)_3(C_2H_3O_2)_9 \cdot 5H_2O$ should form.

Demonstration 5: Reaction of Sodium with Water—Reducing the Rate of Reaction

Purpose

To use an alternate method to demonstrate the reactivity of sodium with water. This demonstration gives students more time to observe the reaction.

Materials

1 beaker, 250-mL or 400-mL, *or*
1 test-tube, 25 x 125-mm or taller, Pyrex
Sodium metal, small piece
Water containing a few drops of phenolphthalein solution (for preparation see *Demonstration 2*)
Kerosene or an equivalent hydrocarbon

Safety

With sodium and all alkali metals, always make sure you have a clean piece of the metal. Oxide coating on the metal surface may result in violent expulsion of the lump from the reacting vessel, with accompanying splashing out of the caustic hydroxide solution. Since kerosene is flammable, **do not** test for hydrogen gas with a lighted splint or other source of flame!

Procedure

1. Add a 3-5 cm layer water and 1-2 drops phenolphthalein to beaker.
2. Add a 3-5 cm layer kerosene to the vessel.
3. Carefully drop in a small piece of sodium.

Discussion

Since the relative densities are water > sodium > kerosene, the sodium will float at the interface between the two immiscible liquids. Reaction of the sodium with water will produce hydrogen gas, which will lift the sodium up into the kerosene layer where the reaction will cease. As the hydrogen gas is evolved the sodium will fall again to the interface and the entire cycle will be repeated. The cycles of reaction can thus be extended to several minutes. This approach increases the observing time for students, and also reduces the intensity of the reaction between the sodium and water.

Reaction: $2Na(s) + 2H_2O(l) \rightarrow H_2(g) + 2NaOH(aq)$

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Why are alkali metals never found in elemental form in nature? *[The alkali metals are extremely reactive. More specifically, they react with water, oxygen, the halogens, and in more complex reaction sequences form hydrogen carbonates, carbonates, and other compounds.]*
2. How are alkali metals prepared or obtained? *[The alkali metals can be prepared by electrolysis of their molten salts:*

$$2MX(s) \rightarrow 2MX(l) \rightarrow 2M(l) + X_2(g)$$

3. Why are many alkali metal compounds soluble in water? *[Solution of a solid is a three-step process. In the first step, ions in the crystal lattice have to be separated. Since alkali metal ions have large ionic radii and only a +1 charge, the coulombic attractive forces in the crystal lattices are weak and more readily separated.]*
4. How are the physical and chemical properties of alkali metals related to their electronic structure? *[The alkali metal atoms are large and have only one valence electron. Their ions are also large and have a charge of +1. These characteristics lead to relatively weak attractions between the atoms in the alkali metals and the alkali metal ions in their compounds. Alkali metals more readily lose their electrons than other elements, leading to high chemical reactivity with elements and compounds.]*
5. How do physical and chemical properties of alkali metals vary in the group? *[As the atoms increase in size as one goes down the column in the Periodic Table, the attractive forces decrease and in general so do the numeric values of properties such as melting point, boiling point, hardness. Correspondingly, the decreasing first ionization energy trend from Li to Cs leads to increased chemical reactivity going down the alkali metal family.]*
6. Why are reactions involving elemental alkali metals oxidation-reduction reactions? *[Reactions involving the chemical elements are all oxidation-reduction reactions. In the case of alkali metals the general reaction is* $M^\circ \rightarrow M^+ + e^-$.*]*
7. Why are reactions involving alkali metal compounds rarely oxidation-reduction reactions? *[Alkali metal compounds contain large* M^+ *ions, which add electrons with difficulty to form the elements.* $e^- + M^+ \rightleftharpoons M$.*]*
8. Why are reactions involving compounds of alkali metals almost all acid-base, precipitation or redox reactions of the anionic constituents? *[The large alkali metal +1 ions have weak coulombic electrical force fields and only weakly attract anions. Hence they form few precipitates. The large hydrated alkali metal ions have a very low tendency to function as acids by donating protons (from the* H_2O *ligands). Since alkali metal ions (see #7) have little tendency to gain electrons they undergo few oxidation-reduction reactions. Instead the reactions of the compounds involve the anions. For example:*

Precipitation $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$

Acid-base $2NaOH(aq) + H_2SO_4(aq) \rightarrow 2H_2O(l) + Na_2SO_4(aq)$
$Na_2CO_3(aq) + 2HCl(aq) \rightarrow CO_2(g) + H_2O(l) + 2NaCl(aq)$

Redox $5Na_2C_2O_4(aq) + 2KMnO_4(aq) + 8H_2SO_4(aq) \rightarrow 10CO_2(g) + 2MnSO_4(aq) + K_2SO_4(aq) + 5Na_2SO_4(aq) + 8H_2O(l)$*]*

9. What is the crystal structure of alkali metals? *[Alkali metals crystallize in body-centered cubic lattices in which only 68% of the unit cell is occupied by cations in the unit cell.]*
10. Is there a relationship between physical properties of alkali metals and crystal structures? Explain. *[As indicated in Question 5, the attractive forces among atoms in alkali metals decrease going down the column resulting in a decrease in the magnitude of all properties related to coulombic attractive forces.]*
11. Why do alkali metals have such low ionization energies? Explain using Coulomb's law. *[Alkali metals have large atomic radii and as a consequence the outer valence electron is only weakly attracted by the core charge of +1. These attractive forces decrease from* Li^+ *to* Cs^+ *as the ionic radii increase. Since in the Coulomb's Law equation (see Language of Chemistry), atomic radius is in the denominator, increasing radius leads to decreased attractive force for the valence electron.]*

Counterintuitive Examples and Discrepant Events

In the halogen family the largest atom, At, and the molecule At_2 are the least reactive. The largest atom in the alkali metal family, Cs, is the most reactive.

Pictures in the Mind

(See *Transparency Master* in the *Appendix*: Cross Section of Atoms and Ions.)

TIPS FOR THE TEACHER

Language of Chemistry

NOTE: Many of the terms used in this module are defined elsewhere.

alkali metals family of elements characterized by their vigorous reaction with water. The elements in this family are lithium, sodium, potassium, rubidium, cesium and francium.

anion negatively charged ion.

atomic radius one half the distance between nuclei of two adjacent atoms of the same element.

cation positively charged ion.

charge density charge on an ion divided by its surface area.

Coulomb's Law relationship between electrical forces, charges and distance: the electrical force between two charged objects varies directly as the product of the charges and inversely as the square of the distance between them $[F = k \times (q^{+} \times q^{-}) / r^{2}]$.

critical pressure minimum pressure that must be applied to bring about liquefaction at the critical temperature.

critical temperature temperature above which a gas will not liquefy.

crystal lattice energy energy required to completely separate one mole of a solid ionic compound into gaseous ions.

electrolysis process, involving either the molten state or an electrolytic solution, by which compounds are decomposed electrically.

electron affinity energy associated with the gain of an electron by a neutral gaseous atom.

electronegativity measure of electron attracting power of an atom; metals have low electronegativities, nonmetals have high electronegativities.

hydration energy energy associated with dissolving gaseous ions in water, usually expressed per mole of ions.

ion charged atom or group of atoms, formed when the atom or group of atoms loses or gains electrons.

ionic radius radius of a spherical ion; it is the radius associated with an element in its ionic compounds.

ionization energy amount of energy needed to remove a single electron from a neutral, isolated (gaseous) atom.

lattice geometric model showing the regular arrangement of atoms or ions in a crystalline solid.

metal one of a group of substances characterized by luster, malleability, ductility, and good electrical and heat conductivity; metals tend to form positive ions in ionic compounds; elements that are metals are located on the left side of the Periodic Table.

ore mineral deposit containing sufficiently high concentration to allow economical recovery of a desired metal.

oxidation number a number assigned to an atom in a neutral molecule or ion to reflect its state of oxidation.

photoelectric cell a chemical cell which requires for its operation the ejection of electrons from specific metal atoms when exposed to light.

second ionization energy amount of energy needed to remove a second electron after a single electron has already been removed from a neutral, isolated ion (gaseous) atom.

unit cell smallest unit of a crystal that, if repeated indefinitely, could generate the whole crystal.

Pattern Recognition

1. Figure 2 gives crystal ionic radii in picometers, as measured by X-ray crystallography. (1 pm = 1×10^{-12} m)

Li^+	60	F^-	136
Na^+	95	Cl^-	181
K^+	133	Br^-	195
Rb^+	148	I^-	216
Cs^+	169		

Figure 2. Crystal ionic radii (pm).

a. Considering chemical periodicity, what trend would you predict for the interionic distances for the fluorides of the alkali metals? For the cesium salts of the four halide ions? (The interionic distance is the distance between the centers of the two ions.) *[Interionic distances will increase as you go down the alkali metal family of fluoride compounds.]*

b. Calculate the interionic distances between the centers of all combinations of halide and alkali metal ions. *[Calculated interionic distances of all combinations of alkali metal and halogen ionic compounds:*

	Li^+	Na^+	K^+	Rb^+	Cs^+
F^-	196	231	269	284	305
Cl^-	241	276	314	329	350
Br^-	255	290	328	343	364
I^-	276	311	349	364	385

Figure 3. Calculated interionic distances.

2. Use the results of your calculations from Problem 1b above and Pauling's values for interionic distances in Figure 4 to compare calculated values of ionic radii of alkali and halide ions to the actual values of interionic distances between centers of alkali metal and halide ions measured from X-ray crystallography data. *[Calculated values are generally lower than Pauling values.]*

Cation	Li^+	Na^+	K^+	Rb^+	Cs^+
F^-	201	231	266	282	300
Cl^-	257	281	314	328	356
Br^-	275	298	329	343	371
I^-	302	323	353	366	382

Figure 4. Calculated interionic distances between center of halide and alkali metal ions (pm).

3. Consider the forces of attraction between alkali metal ions and halide ions in terms of their size and charge (Figure 2). Which pair of ions would form an ionic crystal with the greatest crystal lattice energy? Least crystal lattice energy? *[Lithium fluoride, with the two smallest ions, would have the greatest crystal lattice energy because they will be held together most tightly. Cesium iodide, with the largest ion sizes, should have the least crystal lattice energy. See Figure 5.]*

Cation	Li^+	Na^+	K^+	Rb^+	Cs^+
F^-	1034	914	812	780	744
Cl^-	840	770	701	682	630
Br^-	781	728	671	654	613
I^-	718	681	632	617	585

Figure 5. Crystal lattice energies for alkali halides (kJ/mol).

4. Using Figures 4 and 5 and Coulomb's Law, explain whether trends in crystal lattice energies in Figure 5 appear consistent with the interionic distances in Figure 4. *[The force of attraction between ions increases with increasing charge on the ions and with decreasing size of ions. Since all alkali metal ion pairings with halide ions are identical with respect to charge, size is the determining factor. Therefore the combination involving smallest ion sizes (Li^+ and F^-) should have the greatest attraction and hence the greatest crystal lattice energy (1034 kJ/mol). On this basis Cs^+ and I^-, the largest ions, should have the smallest lattice energy (585 kJ/mol). Trends in crystal lattice energy seem consistent with interionic distances in general;* e.g., *ion combinations with similar interionic distances have similar crystal lattice energies. (For examples, see RbF and LiBr, 282 and 275 pm interionic distances and 780 and 781 kJ/mol lattice energies.)]*

5. The Table of Properties of Alkali Metals in the *Appendix* summarizes an extensive set of properties for the alkali metal elements. A useful procedure is to take the properties one by one and ask students to predict the trend in each property going down the group. Additionally, one can query students as to the relative values for a given property as they relate to properties already covered. The overall objective is to underscore patterns of behavior for the properties of a family of elements.

   A possible sequence of discussion questions follows. While displaying on the overhead projector the data needed for each question, you can uncover the table of properties sequentially.

a. On the basis of atomic number of the alkali metal elements, write the electron configurations. (Note similarities as well as differences.)

Element	*Li*	*Na*	*K*	*Rb*	*Cs*	*Fr*
Elect. Struct.	*[He]*$2s^1$	*[Ne]*$3s^1$	*[Ar]*$4s^1$	*[Kr]*$5s^1$	*[Xe]*$6s^1$	*[Rn]*$7s^1$

[Similarity: All have the same outer energy level electron configuration (ns^1*).*

Difference: Each has its own inner core of filled energy levels (its own noble gas configuration).]

b. What valence (oxidation) state do you predict for the alkali metals? Why? *[*M^+ *oxidation state because each has only one electron in its outer energy level to lose.]*

c. How do the sizes of the atoms change with increase in atomic number? Explain. *[As atomic number increases, size of the alkali metal atom will increase also due to increasing number of filled energy levels of electrons within each atom.]*

d. Which alkali metal has the lowest (highest) ionization energy and loses its electrons most (least) readily? Explain. How does the first ionization energy value change with increasing atomic number? Why? *[Francium has the lowest ionization energy and will lose electrons most easily. This is due to its large atomic size and its relative inability to hold on to its outer electron. Lithium has the highest ionization energy and loses electrons least easily due to its small atomic size and its relative ability to hold on to its outer electron. First ionization energy decreases with increasing atomic number due to the increasing atomic size.]*

e. What can you predict about the second ionization energies of the alkali metals? Would you expect this family of elements to form compounds with 2+ ions? Explain in terms of electronic structures. *[Second ionization energies of alkali metals will be large because, in order to remove a second electron, one must attack the next innermost energy level, which is filled in each case. Since this is so difficult a task, compounds with 2+ ions are not likely to be stable.]*

f. How do the ionic radii and atomic radii of alkali metal elements compare? Explain. *[Ionic radii of alkali metals are all smaller than their respective atomic radii, since alkali metal ions form by removal of the outermost electron which in effect removes the outside energy level and decreases the size of the remaining ion.]*

g. The properties of an ion depend on its charge, radius, and inner electronic structure. How does the attractive coulombic force of an ion change with decreasing size? Increasing charge? *[As the ion gets smaller, the force of attraction becomes larger (as the square of the distance); as the ionic charge increases, the force of attraction becomes larger also.]*

h. List the alkali metal ions in order of increasing electron affinity. *[Electron affinities: Fr < Cs < Rb < K < Na < Li]*

i. How does electronegativity of the alkali metals change going from Cs to Li? *[Electronegativity increases from Cs to Li.]*

j. In view of the decrease in attractive forces in the metallic lattices going from Li to Cs, predict how the following properties change in going down the column of alkali metals: (1) melting point, (2) boiling point, (3) heat of fusion, (4) heat of vaporization, (5) heat of atomization, (6) hardness, (7) critical pressure. *[(1) melting points should decrease, (2) boiling points decrease, (3) heats of fusion decrease, (4) heats of vaporization decrease, (5) heats of atomization decrease, (6) hardness decreases, (7) critical pressure decreases.]*

Optional Questions

k. Discuss the trend in density with respect to trends in metallic radius and atomic mass. *[Density generally increases, meaning that the atomic mass must increase at a greater rate than does the volume.]*

l. Discuss the relationship between trends in heat of fusion, metallic radius, and melting point. *[As one goes down the Periodic Table, the metallic radius of alkali metals increases. The increase in radius results in much decreased attractive forces between atoms within the lattice structure, resulting in a decrease in heat of fusion and melting points because less heat is needed to break apart the solid lattice among the larger alkali metals.]*

m. Discuss the relationship between trends in heat of vaporization, metallic radius, and boiling point. *[As one goes down the Periodic Table, the metallic radius of alkali metals increases. The increase in radius results in much decreased attractive forces between atoms within the liquid metals, resulting in a decrease in heat of vaporization and boiling points because less heat is needed to separate atoms from the liquid state within the larger alkali metals.]*

n. The "ionic potential" combines the properties of both charge and ionic radius into one numerical value. For Li^+

$$\text{Ionic potential} = \frac{+1}{0.060\ \text{nm}} = 16.7\ \text{nm}^{-1}$$

What can you conclude from the numerical values of the ionic potentials, going from Li^+ to Cs^+? *[Ionic potential decrease going from Li^+ to Cs^+. Since the ionic charge is +1 for all alkali metals, this decrease in ionic potential must be due to the increasing ionic radius (in the denominator of the term).]*

o. Consider the surface charge on the alkali metal cations and the resulting difference in attraction for anions.

$$\text{Charge density} = \frac{\text{Charge on cation}}{\text{Surface area of cation } (4\pi r^2)}$$

$$\text{Charge density for } Li^+ = \frac{-1}{4 \times 3.14 \times (0.060)^2} = 22\ \text{nm}^{-2}$$

Compare the charge density of Li and Cs. Which ion will exert more attractive force on nearby anions? Why? *[Lithium ion will exert more attractive force on nearby anions because it has a charge density of 22 nm^{-2}* vs. *2.8 nm^{-2} for cesium. The greater charge density for lithium indicates a greater charge per surface area ratio that will result in greater attractive forces (for anions).]*

p. Remembering the charge density of Li^+ and Cs^+, predict their relative abilities to attract water molecules in solution. What effect would this have on the trend in the size of hydrated alkali metal ions? (See the table of data to check your predictions.) *[Lithium should have a greater attractive force for water molecules than cesium, resulting in the hydrated lithium ion's being larger than the hydrated cesium ion. Lithium's hydrated ion is the largest of all the alkali metals.]*

q. Now think about the rate of movement of the alkali metal ions during electrolysis. Which of the ions will move with the greatest velocity? Did you expect it to be Li? *[Cesium will move fastest since it is the smallest hydrated ion. Lithium would be the most likely choice since the lithium ion is the smallest of the alkali metal ions, but when hydrated, it becomes the largest and is thus slowed down in solution.]*

r. What did you notice about the crystal structure of alkali metals? Consider the metallic crystals as a lattice consisting of M^+ ions and an electron gas. How would you expect the attractive forces within the metal to change from Li to Cs? *[All alkali metal crystal structures are body-centered cubic. The attractive forces between the* M^+ *ions and the electron gas cloud around them should decrease as one goes down the Periodic Table.]*

s. How would you expect the critical temperature and critical pressure to vary from Li to Cs? *[Critical temperature* (T_c) *and critical pressure* (P_c) *should decrease as you go from Li to Cs. Temperature can be viewed as the separating force that is responsible for producing isolated gaseous atoms from the molten metal. Coulombic forces within the liquid are the attractive forces that are responsible for holding atoms together in the liquid (or solid) state. At the critical temperature, these two forces, separation and attraction, are equal. Thus as metallic radius increases from lithium to francium and resulting coulombic forces of attraction decrease, less separating force is needed to produce gaseous atoms and therefore* T_c *will be lower. Since critical pressure is the pressure needed for liquefaction to occur at the critical temperature, a lower temperature will require less pressure to bring about liquefaction.]*

t. The alkali metals having an unpaired electron can form covalently bonded M:M, M_2, molecules. Predict the trend in energies needed to break the bonds in one mole of $M_2(g)$ from Li to Cs. *[The energy required to break the M:M bond from Li to Cs should decrease since the atoms get larger and the electrons involved in the bond are farther from each nucleus in the Cs:Cs bond than they are in the Li:Li bond.]*

u. What do you notice about thermal conductivity of alkali metals? *[More data is needed (3 are blank), but it seems that all the alkali metals are good conductors of thermal energy.]*

v. How would you expect the entropy (degree of disorder) of the solid alkali metals to change going down the Periodic Table? *[Entropy should increase as the atomic size increases because lattice arrangements would allow for more randomness of the electrons within the lattice.]*

6. Electronic Structure and Chemical Periodicity

General Format:

l							
f							
d							
p							
s							
n	*1*	*2*	*3*	*4*	*5*	*6*	*7*

In the horizontal direction the entries correspond to the principal quantum number, n, and to the period (horizontal row) in the Periodic Table. The vertical direction relates to the azimuthal quantum number, l, and lists the number of electrons in the sublevels (*s*, *p*, *d*, and *f*) of each main energy level. As one goes down a family (group) of elements, electrons are found in one additional main energy level. Except for the first element in each family, the electron arrangement in the outer two levels remains the same. If the atomic number increases by 8, a new column of 2,6 is added; if it increases by 18 in addition to the new column of 2,6, ten (10) electrons are added to the third from the outermost column; whereas, with an increase of 32 in the atomic number, the addition adds a group of 14 to the fourth outermost column. Going down the alkali metal column, additions are made internally, but the outermost energy level stays the same; *e.g.*,

Li

l		
s	2	1
n	*1*	*2*

Na

l			
p		6	
s	2	2	1
n	*1*	*2*	*3*

K

l				
p		6	6	
s	2	2	2	1
n	*1*	*2*	*3*	*4*

Rb

l					
d			10		
p		6	6	6	
s	2	2	2	2	1
n	*1*	*2*	*3*	*4*	*5*

Cs

l						
d			10	10		
p		6	6	6	6	
s	2	2	2	2	2	1
n	*1*	*2*	*3*	*4*	*5*	*6*

Fr

l							
f				14			
d			10	10	10		
p		6	6	6	6	6	
s	2	2	2	2	2	2	1
n	*1*	*2*	*3*	*4*	*5*	*6*	*7*

Going across a period; *e.g.*, Na to Ar, no additional vertical columns are added, but the outermost energy level increases one electron at a time; *e.g.*,

l								
p			1	2	3	4	5	6
s	1	2	2	2	2	2	2	2
	Na	Mg	Al	Si	P	S	Cl	Ar

Reference: Andrews and Kokes, *Fundamentals of Chemistry.*

Common Student Misconceptions

1. **"The alkali metals easily ('love to') lose one electron each (to achieve noble gas electron configurations)."**

   All alkali metals require energy to be ionized. Their first ionization energies range from +375 kJ/mol to +520 kJ/mol. This is hardly easily losing electrons! Still, of all groups of metallic elements, alkali metals lose electrons most readily.

2. **"Diagrams in some texts show graphs similar to the enthalpy diagram in Figure 6 as being exothermic overall, when they attempt to explain the E° value trends of the alkali metals."**

   As a careful examination of the diagram below reveals, the overall process is endothermic (*i.e.*, $Li(s) \rightarrow Li^+(aq) + 1\ e^-$, $\Delta H = +167$ kJ). The exothermicity actually results from formation of OH^- ion and H_2 gas, rather than the ionization steps shown in Figure 6.

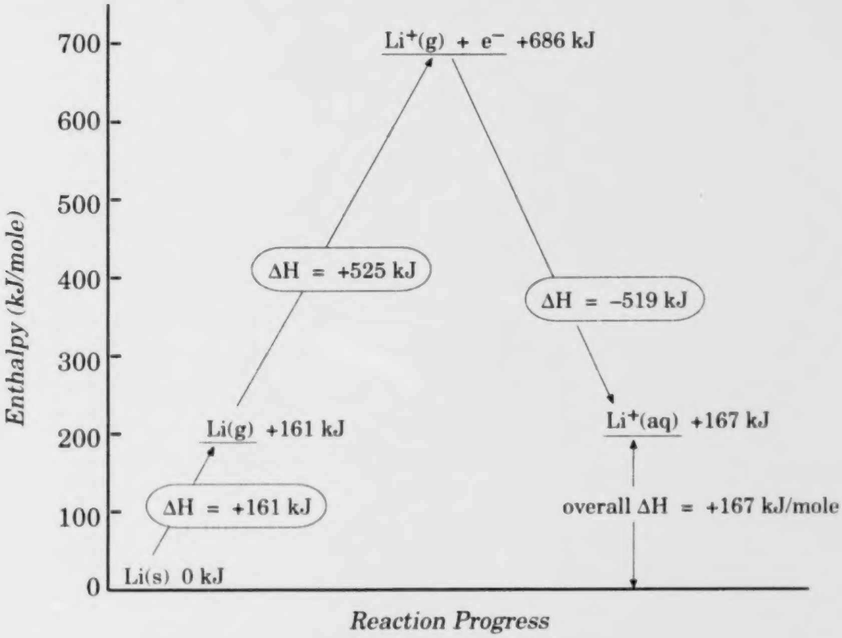


Figure 6. Enthalpy diagram for the ionization of Li.

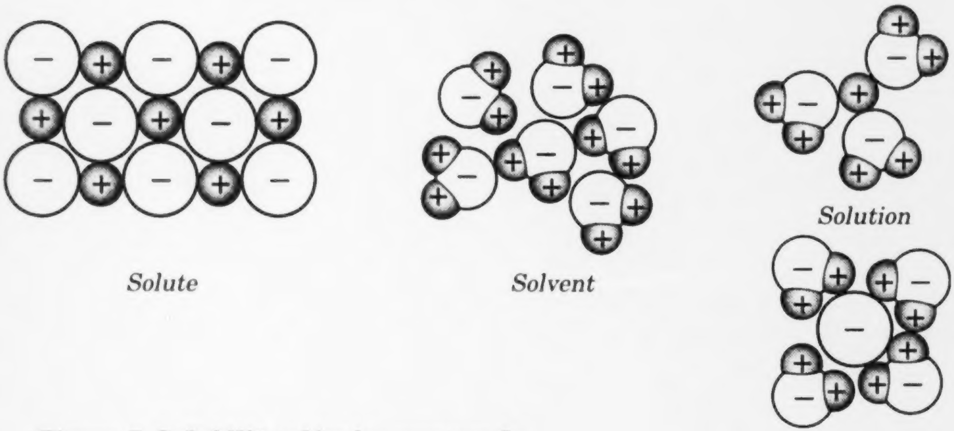
3. **"Many consumer products contain sodium and potassium."**

   Although the ingredients list says "sodium" or "potassium," it is the Na^+ and K^+ ions that are present. The elements and ions have very different properties! Na and K (metals) are toxic, but Na^+ and K^+ (ions) are essential for life. It is important that students recognize the difference between free elements and ions.

4. **"Salts of the heavier members of a group are always more soluble in water than salts of the lighter elements of that group."**

   The solubility of ionic compounds can be visualized as consisting of three steps:

Step 1:	breaking the crystal lattice (requires energy)	$\Delta H = +$
Step 2:	formation of a hole in the solvent into which ions will fit (requires energy)	$\Delta H = +$
Step 3:	hydration of ions (releases energy)	$\Delta H = -$


Solute

Solvent

Solution

Figure 7. Solubility of ionic compounds.

When discussing solubility trends in terms of chemical periodicity, many teachers try to explain changes in terms of electrical coulombic attractions among the ions in the crystal lattice. Thus in the series $Be(OH)_2$, $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$, and $Ba(OH)_2$, the last compound is most soluble because the larger barium ion has the weakest attraction for the hydroxide ion in the crystal lattice of all the alkaline earth ions in this series. Similarly CsOH is more soluble and more basic because it is the largest ion of all the alkali metals. However, this explanation fails when trying to explain the decreasing solubility in the alkaline earth sulfate series, $BeSO_4$, $MgSO_4$, $CaSO_4$, $SrSO_4$, and $BaSO_4$. The misconceptions arise for at least two reasons:

a. You must consider the three steps in the solution mechanism and the overall ΔH of solution. Major differences in solubility of ionic solids will arise primarily from Steps 1 and 3, since the formation of the hole in the solvent, Step 2, is approximately the same for most ions.

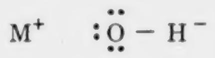
b. Whether or not an overall process is (thermodynamically) possible depends upon the free energy change ($\Delta G = \Delta H - T\Delta S$) that includes both enthalpy and entropy effects. Because dissolved species are more random than the highly ordered crystal lattice, the entropy effect always favors dissolving.

> For solids that dissolve by absorbing heat, such as ammonium chloride, the enthalpy effect opposes dissolving. Thus solubility can frequently be pictured as a tug of war between these two effects. Since free energy change, ΔG, must be negative for a process to occur spontaneously, and $\Delta G = \Delta H - T\Delta S$, where T is Kelvin temperature and ΔS is the entropy change, a positive ΔH can be overcome only if T is sufficiently large.

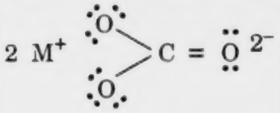
Clearly +1 and –1 ions do form lattices with relatively weak lattice energies, but they also tend to have the weakest attractions for water molecules in the hydration step because of the large size and small charge.

Problem Solving

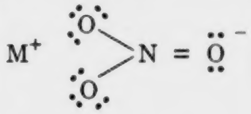
1. Remembering that alkali metals constitute a family of chemical elements, as do the halogens (F, Cl, Br, I), and given NaCl as the formula of sodium chloride, predict formulas for each of the following compounds:
   - a. Sodium fluoride *[NaF]*
   - b. Potassium iodide *[KI]*
   - c. Lithium chloride *[LiCl]*
   - d. Cesium iodide *[CsI]*
   - e. Rubidium bromide *[RbBr]*
   - f. Any alkali metal, M, with any halogen, X *[MX]*
2. Examine the general structures for hydroxides, carbonates, and nitrates shown:



Hydroxide



Carbonate



Nitrate

On the basis of these structures and the differences in physical properties of the metallic ions, explain why the following products were formed in these reactions:

a. $2\,LiOH(s) + Heat \rightarrow Li_2O(s) + H_2O(g)$

b. $Li_2CO_3(s) + Heat \rightarrow Li_2O(s) + CO_2(g)$

c. $4\,LiNO_3 + Heat \rightarrow 2\,Li_2O(s) + 4\,NO_2(g) + O_2(g)$

d. $NaOH(s) + Heat \rightarrow NaOH(l)$

e. $Na_2CO_3(s) + Heat \rightarrow Na_2CO_3(l)$

f. $2\,NaNO_3(s) + Heat \rightarrow 2\,NaNO_2(s) + O_2(g)$

Note that other alkali metals form products similar to those of sodium, while magnesium, an alkaline earth element, forms products similar to those of lithium. *[When heated the +1 alkali metal ion with the highest surface charge density, the lithium ion will attract the O^{2-} in the OH^-, CO_3^{2-} or NO_3^- most strongly to form the oxide Li_2O.]*

3. Predict whether calcium will react more like lithium or sodium in similar reactions. *[All are related to sizes of and charges on metallic ions (density of positive charge on the ion's surface). The smaller the size and the greater the charge, the more strongly the metallic ion attracts the oxygen ion from the anion. Calcium, with greater nuclear charge and slightly smaller ion size, will react more like lithium than like sodium.]*

4. On the basis of oxidation states, the Periodic Table, and the fundamental classes of chemical reactions, predict the products formed in the following reactions: [Products are given in brackets. These equations are not balanced.]

a. $Na(s) + Cl_2(g) \xrightarrow{heat}$ *[NaCl]*

b. $Li(s) + O_2(g) \xrightarrow{heat}$ *[Li_2O Note: Other alkali metals do not react with oxygen directly to produce simple oxides.]*

c. $K(s) + H_2(g) \xrightarrow{heat}$ *[KH]*

d. $Rb(s) + Br_2(g) \xrightarrow{heat}$ *[RbBr]*

e. $NaCl(l) \xrightarrow{heat}$ *[$Na + Cl_2$]*

f. $KCl(s) + Na(s) \rightarrow$ *[$K + NaCl$]*

g. $Na(s) + HOH(l) \rightarrow$ *[$NaOH + H_2$]*

h. $Li(s) + N_2(g) \rightarrow$ *[Li_3N]*

i. $NaH(s) + HOH(l) \rightarrow$ *[$NaOH + H_2$]*

j. $Na_2O(s) + HOH(l) \rightarrow$ *[NaOH]*

k. $NaOH(s) + H_2SO_4(aq) \rightarrow$ *[$NaHSO_4$ (or Na_2SO_4) + H_2O]*

l. $NaOH(s) + CO_2(g) \rightarrow$ *[$NaHCO_3$ (or Na_2CO_3) + H_2O]*

m. $NaCl(s) + H_2SO_4(aq) \rightarrow$ *[$NaHSO_4$ (or Na_2SO_4) + HCl]*

n. $K_2O(s) + HCl(aq) \rightarrow$ *[$KCl + H_2O$]*

o. $KOH(aq) + HNO_3(aq) \rightarrow$ *[$KNO_3 + H_2O$]*

p. $2NaCl(s) + 2HOH(l) \xrightarrow{electrolysis}$ *$Cl_2 + H_2 + NaOH$]*

q. $Cs(s) + N_2(g) \xrightarrow{heat}$ *[No reaction]*

5. All the alkali metals crystallize in body-centered cubic lattices. How many atoms are there in a unit cell (Figure 8)?

Figure 8 illustrates a body-centered cubic lattice. *[The number of particles that "belong" to a given unit cell consists of all interior particles, which belong exclusively to that unit cell plus fractions of those that are shared by two or more unit cells. Since each particle at a corner belongs to eight unit cells (see corner A), only 1/8 of a corner particle belongs to the unit cell. There are eight corners; thus, each unit cell has 1/8 x 8 = 1 particle in addition to interior particles. Since alkali metals crystallize in body-centered cubic (BCC) cells, their unit cells contain 2 atoms (1 in the center and 1/8 x 8 = 1 at the corners).]*

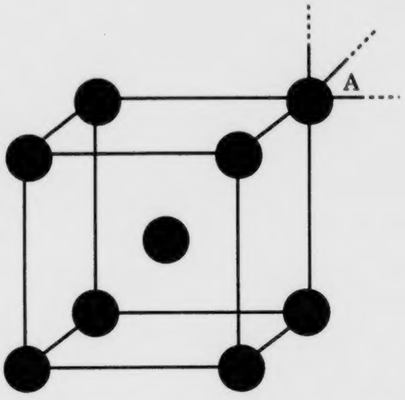


Figure 8. Body-centered cubic lattice.

6. The vapors above heated liquid alkali metals contain about 1% diatomic molecules. The equation below describes the reaction.

$$2M(g) \rightleftharpoons M_2(g)$$

Write the Lewis structure for the $Na_2(g)$ molecule. *[Na:Na]*

7. Predict the products you would obtain by the electrolysis of molten lithium hydride, LiH. *[Li + H_2]*

8. Balance each of the following oxidation reduction equations by a systematic method in which the number of electrons lost in the oxidation half-reaction is equal to the number of electrons gained in the reduction half. Identify the substance oxidized, the substance reduced, the oxidizing agent and the reducing agent.

   a. $NaCl + HOH \rightarrow H_2 + Cl_2 + NaOH$

   [$2NaCl + 2HOH \rightarrow H_2 + Cl_2 + 2NaOH$; Cl is oxidized, H is reduced, HOH is the oxidizing agent, NaCl is the reducing agent.]

   b. $Ca + CsCl \rightarrow Cs + CaCl_2$

   [$Ca + 2CsCl \rightarrow 2Cs + CaCl_2$; Ca is oxidized, Cs is reduced, CsCl is the oxidizing agent, Ca is the reducing agent.]

9. The ionic mobility (molar ionic conductance) of very dilute aqueous solutions of alkali metals (at 18 °C) are given in Table of Properties of Alkali Metals in the *Appendix*. Explain why aqueous Cs^+ ion moves the fastest when electrolyzed. *[The hydrated Li^+ is larger than the less hydrated Cs^+.]*

10. Cesium is found in the mineral pollucite, $Cs_4Al_4Si_9O_{26} \cdot H_2O$, on the island of Elba. Calculate the percent by mass of cesium in this mineral.

4 Cs	*4 moles x 132.9 g / mol*	=	*531.6 g*
4 Al	*4 moles x 27.0 g / mol*	=	*108.0 g*
9 Si	*9 moles x 28.1 g / mol*	=	*252.9 g*
26 O	*26 moles x 16.0 g / mol*	=	*416.0 g*
1 H_2O	*1 mole x 18.0 g / mol*	=	*18.0 g*
			1326.5 g / mol

% Cs = (531.6 g / 1326.5 g) x 100 = 40.1 % Cs

11. The gaseous alkali metal atoms tend to form diatomic molecules:

$$2M(g) \rightarrow M_2(g)$$

Predict the value for the heat of formation of Rb_2 from the following data:

Li_2 $\Delta H_f^o = -113$ kJ

Na_2 $\Delta H_f^o = -77.0$ kJ

K_2 $\Delta H_f^o = -52.7$ kJ

Cs_2 $\Delta H_f^o = -43.5$ kJ

The experimental value for Rb_2 is –47.3 kJ. *[(–52.7 -43.5) / 2 = –48.1 kJ, by interpolating the K_2 and Cs_2 values.]*

Why would you expect alkali metals to form M_2 molecules? *[Atoms with unpaired electrons tend to react by pairing or transferring electrons.]*

12. Chemists have observed that the elemental pairs Li-Mg, Be-Al, and B-Si have similar chemical properties, which is known as the diagonal relationship. Calculate the ratio of the charge on the ion:ionic radius (called the ionic potential) for the following ions, and discuss your results related to diagonal relationship.

Element	Li	Na	Be	Mg	Ca	B	Al	C	Si
Charge	+1	+1	+2	+2	+2	+3	+3	+4	+4
Ionic radius, (pm)	60	97	35	66	99	23	51	16	42
Ionic potential	*0.017*	*0.010*	*0.057*	*0.030*	*0.020*	*0.130*	*0.059*	*0.250*	*0.095*

The ionic potentials of Be-Al and B-Si are fairly close, but that of Li is about half that of Mg.

Decision Making

1. How can students systematically predict whether or not a solid $A_m^{n+}B_n^{m-}$ will be soluble in water? A convenient approach is to apply the following steps *in order*, based upon the two ions composing the solid, stopping at the step in which one of the ions first appears.

Step 1. If the cation is H^+, Na^+, K^+, Rb^+, Cs^+, or NH_4^+, the salt is *soluble*.

Step 2. If the anion is NO_3^-, NO_2^-, $C_2H_3O_2^-$, ClO_4^-, or HCO_3^-, the salt is *soluble*.

Step 3. If the cation is Ag^+, Hg_2^{2+}, Pb^{2+}, or Tl^+, the salt is *insoluble*.

Step 4. If the anion is Cl^-, Br^-, or I^-, the salt is *soluble*.

Step 5. If the anion is OH^-, O^{2-}, S^{2-}, CO_3^{2-}, or PO_4^{3-}, the salt is *insoluble*.

Step 6. If the anion is SO_4^{2-} and the cation is not Pb^{2+}, Ca^{2+}, Sr^{2+}, or Ba^{2+}, the salt is *soluble*.

On the basis of the above rules, classify the following salts as soluble (S) or insoluble (I).

a.	$AgNO_3$	*[S]*	f.	$Al(NO_3)_3$	*[S]*	k.	$Mg(OH)_2$	*[I]*
b.	$PbSO_4$	*[I]*	g.	$BaCO_3$	*[I]*	l.	$CsNO_2$	*[S]*
c.	Na_3PO_4	*[S]*	h.	SrO	*[I]*	m.	Sb_2S_3	*[I]*
d.	AgI	*[I]*	i.	$Cr_2(SO_4)_3$	*[S]*	n.	$Al(OH)_3$	*[I]*
e.	KCN	*[S]*	j.	$HClO_4$	*[S]*	o.	$Ba_3(PO_4)_2$	*[I]*

For more on the solubility rules, see *Solubility and Precipitation* module.

2. The fluorides, carbonates, and phosphates of all of the alkali metal ions except lithium are soluble in water. Explain lithium's unique solubility properties. *[The crystal lattice energies are larger because of the small size of the lithium ion. See the item on solubility in the* Common Student Misconceptions *section.]*

3. Sodium chloride is typically used on roads in the northeast in winter time to help to melt ice and snow. Unfortunately, the runoff from these roads is a very concentrated salt solution that kills plants growing alongside the road. Students may discuss the pros and cons of this application of salt. Could other compounds be used instead (magnesium chloride, for example)? *[Select salts that provide nutrients needed for plant growth, such as NH_4NO_3, $(NH_4)_2SO_4$,* etc.*]* Students can research and explain to the class the phenomena which account for this process. *[Freezing point depression and cell plasmolysis.]*

Other

Interest Items

1. Sodium is one of the least expensive of all metals. One pound of sodium costs only about 15 cents when purchased in large quantities.
2. It is estimated that there are only 15 g of the element francium in the top kilometer of the lithosphere.
3. An alloy composed of 12% sodium, 47% potassium and 41% cesium has a melting point of −78 °C, lower than any other known alloy. Imagine a metal made of three alkali metals which melts at a temperature far below room temperature!
4. K-40 is a naturally occurring isotope of potassium and makes up 0.0118% of all the potassium in the lithosphere. It is radioactive, decaying with a half-life of 1.2×10^9 yr. Since potassium is an element essential to the life process, and present in all normal diets, what do the above facts imply about potassium's effect on living things? *[All living things contain the radioactive isotope of potassium to the same level as its percentage in the lithosphere. Since this level is so small, and since the half-life of K-40 is so long, the level of radioactivity of the K-40 is very low, and contributes to the background radiation that all things experience.]*

HISTORY: ON THE HUMAN SIDE

As a consequence of their chemical reactivity, alkali metals are not found free in nature. Their preparation awaited the invention of the voltaic pile (battery) by Alessandro Volta in 1799, which provided a technology for adding electrons to the M^+ ions to produce the metals. Another factor in the date of discovery is related to the rank order and mass percent abundance of these elements as ionic compounds in the lithosphere (earth's crust). The rank order and abundances are: Li, 27 (0.0065%); Na, 6 (2.74%); K, 7 (2.47%); Rb, 19 (0.028%); and Cs, 52 (0.00032%). Humphry Davy, in 1807, at the age of 29, prepared metallic potassium by the electrolysis of molten caustic potash (KOH); a few days later he prepared sodium from molten caustic soda (NaOH).

$$4KOH(l) \rightarrow 4K(l) + O_2(g) + 2H_2O(l)$$

$$4NaOH(l) \rightarrow 4Na(l) + O_2(g) + 2H_2O(l)$$

While serving as a laboratory assistant with the famous Swedish chemist, Jons Jacob Berzelius, J. A. Arfvedson in 1817 observed that lithium compounds had properties that were similar to those of sodium and potassium, except for the lower solubilities of the hydroxide and carbonate. Lithium, as an aqueous ion, was isolated by Arfvedson from petalite ($LiAlSi_4O_{10}$), spodumene ($LiAl(SiO_3)_2$, and lepidolite ($K_2Li_3Al_4Si_7O_{21}(OH,F)_3$).

Davy recognized the plant origin of potassium from impure potassium carbonate in wood ashes or "pot ashes" with the symbol for potassium, K, coming from the Latin equivalent, *kalium*. Sodium relates to the plant origin of soda ash with *natrium*, the Latin equivalent, giving the symbol Na. Lithium was chosen by Arfvedson to indicate its origin from stones (Greek, *lithos*). In 1818, Davy prepared metallic lithium by the electrolysis of molten Li_2O.

$$2Li_2O(l) \rightarrow 4Li(l) + O_2(g)$$

Robert Bunsen and Gustav Kirchhoff, using the spectroscope they invented in 1859, discovered two of the alkali metals. In 1860, during a spectroscopic investigation of the brine mineral waters in a spa in Durkheim, they observed two bright double blue lines. The name cesium derives from the Latin *caesius*, meaning sky blue. A few months later they viewed the spectrum of a solution obtained by the chemical treatment of lepidolite and noted ruby-red lines. The Latin *rubidus*, meaning "deepest red," gave rise to the name, rubidium. In 1934, Marguerite Perey, a French radiochemist, identified element 87 as a decay product of actinium-227, which in turn ultimately comes from uranium-235. In honor of her native country she named the element francium.

Davy, widely known for his invention of the safety lamp, described his discovery of potassium as follows:

> A small piece of pure [caustic] potash [KOH] which had been exposed for a few seconds to the atmosphere, so as to give conducting power to the surface [by attraction of moisture and slight deliquescence], was placed upon an insulated disc of platina [platinum], connected with the negative side of the battery in a state of intense activity; and a platina wire, communicating with the positive side, was brought in contact with the upper surface of the alkali. The potash began to fuse at both its points of electrization. There was a violent effervescence at the upper surface; at the lower, or negative surface, there was no liberation of elastic fluid, but small globules having a high metallic lustre, and being precisely similar in visible characters to quicksilver [mercury], appeared, some of which burnt with explosion and bright flame, as soon as they were formed, and others remained, and were merely tarnished, and finally covered with a white film that formed on their surfaces. These globules, numerous experiments soon showed to be the substance I was in search of, and a peculiar inflammable principle the basis of potash.

Davy named the new element potassium and then by further experiments prepared sodium, calcium, barium, strontium, and magnesium by electrolysis.

Pliny in his writings mentions that the Egyptians made caustic soda [NaOH] by boiling natron [Na_2CO_3] with quicklime [CaO]. Duhamel (1736) and A. Marggraf (1757) showed that potash [K_2CO_3] from wood ashes and soda [Na_2CO_3] from the ashes of marine plants could be distinguished by two tests:

1. Heat on platinum wire: potash gives lilac flame; soda gives a yellow flame.
2. Add platinic chloride to a solution in HCl: potash gives yellow precipitate; soda forms no precipitate.

The Hebrews knew that vinegar effervesced with natural sodium carbonate. In Proverbs 25:20, "... like vinegar poured on soda is one who sings songs to a heavy heart."

Gunpowder was made by the Chinese in about 1150 A.D. for use in making fireworks. This mixture of 75 parts KNO_3, 13 parts carbon, and 12 parts sulfur by mass was reinvented in the western world by Roger Bacon in about 1248 A.D.

Gay-Lussac and Thénard (1808) showed that molten caustic potash (KOH) or caustic soda (NaOH) brought into contact with red-hot iron turnings produced the respective alkali metal as a distillate.

Castner (1886) produced sodium on a large scale by heating NaOH with iron and carbon at a temperature of 1000 °C:

$$6NaOH + 2C \rightarrow 2Na + 3H_2 + 2Na_2CO_3$$

By 1890, Castner developed a large scale electrolytic method for preparing sodium using a cylindrical iron pot with an iron cathode and a nickel anode.

Joseph Black (1728-1799) in his book, *Dissertation on Magnesia* (1754), summarized his research on the chemical nature of the alkalies.

Limestone + Phlogiston → Quicklime (caustic)

By the "principle of causticity" limestone absorbs phlogiston from the fire to produce caustic lime. Soon after this Black showed that in this process there was a loss of mass, and he revised his equation.

Limestone → Quicklime + Fixed air

Our modern interpretation is:

$$CaCO_3 \rightarrow CaO + CO_2$$

Black's other chemical deductions with modern equivalents are:

1. Slaking of quicklime:

   Phlogiston + Quicklime + Water → Slaked quicklime

   $$CaO + HOH \rightarrow Ca(OH)_2$$

2. Caustification of mild alkalis:

   $$\text{Slaked quicklime + Mild alkali} \xrightarrow{\text{boiling}} \text{Caustic alkali}$$

   $$Ca(OH)_2 + K_2CO_3 \rightarrow 2KOH + CaCO_3$$

3. Action of acid on limestone or mild alkali:

   Acid + Limestone → Fixed air

   $$2HCl + CaCO_3 \rightarrow CO_2 + CaCl_2 + H_2O$$

   or

   Acid + Mild alkali → Fixed air

   $$2HCl + K_2CO_3 \rightarrow CO_2 + 2KCl + H_2O$$

4. Precipitation of lime salt by mild alkali:

   Limestone solution + Mild alkali → Limestone + Salt

   $$CaCl_2 + K_2CO_3 \rightarrow CaCO_3 + 2KCl$$

HUMOR: ON THE FUN SIDE

1. **Joke**: Identify the fruit represented by the formula containing one part barium and two parts sodium *[*$Ba(Na)_2$ *— BaNaNa]*
2. **Some rhymes**:
   a. The alkaloid natives of Pollux
   Engage in the strangest of frolics
   They all get their kicks
   Sucking alkali sticks,
   Which, of course, makes them all alkahollux!

   CHEM 13 NEWS, October 1970, p. 195

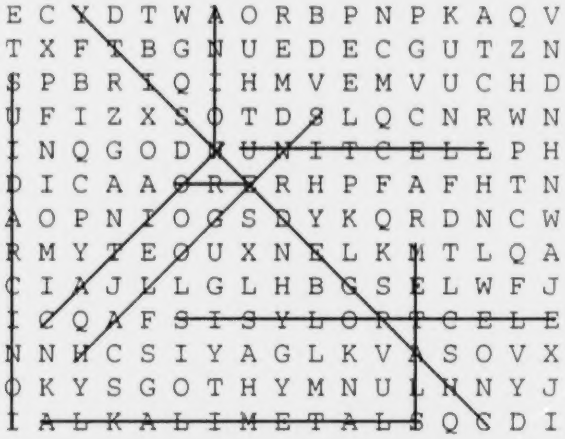
b. Once upon a time
On the border of a flask
Sat a wicked little atom
Who had ne'er done a task.
Did a funny little idiom,
And fell into the sodium.
Had a sad and tragic ending
Once upon a time.

CHEM 13 NEWS, November 1972, p. 516

c. Said a globule of sodium, so sweet
To a droplet of water, petite,
"Let's unite in the air,
Rearrange ourselves there,
And give off a great deal of heat!"

CHEM 13 NEWS, September 1969, p. 10

3. "The Great Sodium Disaster," videotape by Clifford Schrader (see *Media*)
4. Word Search (see *Appendix* for master copy)



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E C Y D T W A O R B P N P K A Q V
T X F T B G N U E D E C G U T Z N
S P B R I Q I H M V E M V U C H D
U F I Z X S O T D S L Q C N R W N
I N Q G O D N U N I T C E L L P H
D I C A A O R E R H P F A F H T N
A O P N I O G S D Y K Q R D N C W
R M Y T E O U X N E L K M T L Q A
C I A J L L G L H B G S E L W F J
I C Q A F S I S Y L O R T C E L E
N N H C S I Y A G L K V A S O V X
O K Y S G O T H Y M N U L H N Y J
I A L K A L I M E T A L S Q C D I
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Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Family of elements characterized by its vigorous reaction with water (2 words).
2. Negatively charged ion.
3. Positively charged ion.
4. Charge on an ion divided by its surface area (2 words).
5. Process by which compounds are decomposed electrically.
6. Family that is very reactive with alkali metals in which binary ionic compounds in a one-to-one ratio are formed.
7. Property of alkali metals that increases from Li^+ to Fr^+ (2 words).
8. Group of elements characterized by luster, malleability, ductility, and good electrical and heat conductivity.
9. Mineral deposit containing sufficiently high concentration to allow practical recovery of a metal.
10. Smallest part of a crystal that, if repeated, could generate the whole crystal (2 words).

Answers: 1. ALKALI METALS 2. ANION 3. CATION 4. CHARGE DENSITY 5. ELECTROLYSIS 6. HALOGENS 7. IONIC RADIUS 8. METALS 9. ORE 10. UNIT CELL

5. See cartoons at end of module.

MEDIA

1. "Chemical Families," videotape of a CHEMStudy film, distributed by Ward's Scientific Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012, Telephone: (716) 359-2502, FAX: (716) 334-6174. Excellent film showing chemical reactions of the alkali metals and lots of other elements. Shows the relationships within families in the Periodic Table. Also available from the same company as a 16 mm film.

2. "The Great Sodium Disaster," by Clifford Schrader. A homemade video of several 1- and 5-pound chunks of sodium and potassium being tossed into an abandoned mine pit. Some great, and some not-so-great photography and sound. Distributed by Weird Science, c/o L. Marek, Naperville High School, Naperville, IL.

3. "Databases in the Classroom: AppleWorks Sampler," 1986 Minnesota Educational Computer Consortium, St. Paul, MN 55165. The program contains a database of "The Elements of the Periodic Table," fully compatible with AppleWorks—very useful to organize large amounts of data by sorting and arranging into meaningful patterns.

4. "MECC Database Sampler," 1985 Minnesota Educational Computer Consortium, St. Paul, MN 55165. This program uses a rudimentary database developed by MECC, called DATAQUEST. It has much of the same data as the AppleWorks Sampler above, but much less ability to manipulate (sort or arrange data). This program may be more available than the AppleWorks Sampler above, since it is part of (and therefore free with) the MECC School Site Licensing Agreement Program, in which many schools nationwide participate. The AppleWorks Sample disk and documentation is sold separately and is not free with the licensing program.

5. "Periodic TableWorks," JCE Software, August 1988. This program is a publication of the Journal of Chemical Education, available from the Project SERAPHIM, Department of Chemistry, University of Wisconsin, 1101 University Avenue, Madison, WI 53706. According to the brochure, this program "converts Apple II computer into an electronic Periodic Table. It highlights elements, groups, or periods; it then displays states, data, or trends of selected properties." It does what it says! It works great for periodicity and trends.

6. "Chemical Families," a CHEM Study film, distributed by Ward's Natural Scientific Establishment. Excellent film showing chemical reactions of the alkali metals and other elements. Shows the relationships within families on the Periodic Table. Also available from the same company as a videotape (VHS).

7. "The Periodic Table" BBC Productions for the Open University, distributed by the Media Guild, 11722 Sorrento Valley Road, Suite E, San Diego, CA 92121-1021. Phone: (619) 755-9191. The film shows good reaction sequences of the alkali metals and halogens. It includes animation and humor.

8. "Alkali Metal Reactions with Chlorine and with Water" Modern Learning Aids CHEM Study Series 092 #4004 This filmloop contains the same close-up sequences as the Chemical Families film. Its advantage over the film is that you can stop the loop to ask questions or slow it down to see frame by

frame (almost). Its disadvantage is that you must explain what students are seeing, as there is no sound.

9. "Physical Properties of the Alkali Metals" Fitch, R. M., and J. T. Fitch, Harper and Row Publishers 04-94872, Harper & Row Publisher, Inc., Keystone Industrial Park, Scranton, PA 18512, FAX: (717) 343-3611.
10. "Chemical Properties of the Alkali Metals" Fitch, R. M., and J. T. Fitch, Harper and Row Publishers 04-94914.
11. "Formation by Electrolysis" Fitch, R. M., and J. T. Fitch, Harper and Row Publishers 04-94455.
12. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396; (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    a. *KC? Discoverer with Knowledgeable Counselor,* by Daniel Cabrol, John W. Moore and Robert C. Rittenhouse. Special Issue 2, for IBM PS/2, PC compatible computers.

    b. *KC? Discoverer: Exploring the Properties of the Chemical Elements,* by Aw Feng and John W. Moore. Vol. I B, No. 1, for IBM PS/2, PC compatible computers.

    c. *KC? Discoverer?,* by Michael Liebl, Vol. IV A, No. 2, for all Apple II computers.

    d. *The Periodic Table Stack,* by Michael Farris. Vol. I C, No. 1, for the Apple Macintosh.

    e. *The Periodic Table (ToolBook),* by Paul F. Schatz, John C. Kotz and John W. Moore, in press, for Windows running on IBM PS/2 PC-compatible computers.

13. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396; (608) 263-2837 (voice) or (608) 262-0381 (FAX).

    For the Apple II computer: AP 207

14. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396; (60) 262-5153 (voice) or (608) 262-0381 (FAX).

    a. "Similarities and Trends in Groups II – Alkali Metals," a chapter on the videodisc *Demonstrations in Organic Chemistry* (double sided, 60 min.), Special Issue 6.

    b. *The Periodic Table Videodisc* (single side, 30 min.). Special Issue 1.

Links/Connections

WITHIN CHEMISTRY

1. Alkali metals provide links to the Periodic Table and periodic properties.
2. Extraction from naturally occurring ores and chemical reactivity of alkali metal elements are both examples of oxidation-reduction reactions.
3. Sodium is obtained by the electrolysis of a mixture of 60% by mass of calcium chloride, and 40% by mass of sodium chloride. The mixture melts at 580 °C, whereas pure sodium chloride melts at 801 °C. This is an example of how the concept of freezing point depression can have practical applications—in this case, a marked reduction in the fuel costs needed to melt the electrolyte. (Both the sodium and calcium are liberated at a stainless steel cathode in the Downs cell and rise through a cooled collecting pipe. The calcium—melting point, 851 °C—solidifies and falls back into the melt while the sodium—melting point, 98 °C—rises and condenses in the collecting chamber.)

   Another example of the use of freezing point depression involves the preparation of lithium. Lithium is prepared from spodumene, $LiAl(SiO_3)_2$:

$$LiAl(SiO_3)_2 \xrightarrow{H_2SO_4} Li_2SO_4 \xrightarrow[250\ ^\circ C]{Na_2CO_3} Li_2CO_3$$

$$Li_2CO_3 \xrightarrow{HCl} LiCl \xrightarrow[\text{electrolysis}]{45\%\ KCl} Li + Cl_2$$

   KCl lowers the melting point of LiCl from 614 °C to 450 °C.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Physics

The bright yellow light of the sodium spectrum is one of the dominant colors in the sun's spectrum (and many street lamps).

Earth Science

1. Deposits of sodium and potassium compounds (*e.g.*, NaCl, KNO_3) are found throughout the world. Large salt deposits are found in Oklahoma, Texas and Kansas.
2. Rivers dissolve and carry 160 million tons of salt to the oceans each year.
3. The sodium chloride from all the oceans would occupy 1.9×10^7 km^3. This volume is equivalent to 150% of the North American continent above sea level. This quantity of solid would form a column with an area of 1 km^2 and a height 47 times the distance to the moon.
4. The Cheshire salt field in Great Britain is 60 km x 24 km with a thickness of 400 m and contains 1×10^{14} kg (1×10^{11} metric tons) of sodium chloride, NaCl.
5. Analysis of mineral water from locations such as Durkheim, Ungemach, and Bourbonne-les-Bains shows Rb^+ and Cs^+ at levels as high as 18.7 mg/L for Rb^+ and 32.5 mg/L for Cs^+.

Biology

1. The analysis of blood serum of all animals indicates 0.022% potassium and 0.32% sodium by mass. The milk of meat eating animals contains equal amounts of the two elements (a 1:1 ratio), whereas in the milk of herbivores and humans, the K to Na ratio is 3.5:1.
2. KI is added to table salt (NaCl) sold in supermarkets to help prevent goiter, a disease associated with iodine deficiency.
3. Marine plants take up potassium ion from the low concentrations in sea water. Potassium salts can be extracted from the ashes of calcined kelp.
4. The perspiration of sheep is so rich in potassium salts that the brown solution obtained from washing the wool in water yields, after evaporation and heating the residue, 5 parts of K_2CO_3 for every 100 parts of wool.
5. The minimum amount of potassium in a fertile soil is 0.01% by mass. An acre of trees removes about 1.5 lbs of K_2CO_3 from the soil each year. Since other growing plants remove even greater amounts, the potassium must be replenished through the addition of fertilizers.
6. Rubidium salts are absorbed from the soil by plants, but cesium salts are vegetable poisons.

TO THE CONTEMPORARY WORLD

Personal

1. Synthetic greases (one component is lithium stearate) are produced using tallow (animal fat), lithium hydroxide, and water, analogous to the production of soaps using potash (KOH).
2. Baking soda, $NaHCO_3$, relieves stomach acidity by the reaction:

   $$HCO_3^-(aq) + H^+(aq) \rightarrow CO_2(g) + H_2O(l)$$

   This reaction is also responsible for leavening of quick breads using baking powder or baking soda and sour milk. (Only the source of H^+ ions is different.)
3. Potash (KOH) is used to make soap, along with animal fat or tallow.
4. The basic mixture used in fireworks is called black powder—a mixture of powdered potassium nitrate (KNO_3) carbon, and sulfur. In place of potassium nitrate, potassium chlorate ($KClO_3$) or potassium perchlorate ($KClO_4$) is often used. To produce the special effects so necessary to pyrotechnic displays, the following compounds are used.

Red flame	$Sr(NO_3)_2$, $SrCO_3$, $LiClO_4$
Green flame	$Ba(NO_3)_2$, $Ba(ClO_4)_2$, $CuCl_2$
Blue flame	$CuCO_3$, $CuSO_4$, CuCl, CuO
Yellow flame	$Na_2C_2O_4$, Na_3AlF_6
Violet flame	CsCl, RbCl, KCl
White flame	Mg, Al
Gold sparks	iron filings
White smoke	KNO_3/S mixture
Colored smokes	KNO_3/S mixture and organic dyes

5. For many years (approximately 1930-1980) the primary industrial use for sodium was in the manufacture of tetraethyl lead, $Pb(C_2H_5)_4$, which was used in most motor fuels to improve octane number and reduce valve knock. Tetraethyl lead does not contain sodium, but the sodium was used to make Na_4Pb, which was reacted with ethyl chloride to make tetraethyl lead.
6. Some common names of alkali metal compounds:

Table salt	$NaCl$
Lye	$NaOH$
Saltpeter	KNO_3
Chile saltpeter	$NaNO_3$
Washing soda	Na_2CO_3
Baking soda	$NaHCO_3$
Borax	$Na_2B_4O_7$
Photographer's "hypo"	$Na_2S_2O_3$
Sylvite	KCl

Community

1. KO_2, potassium superoxide, is used in self-contained breathing apparatus for firefighters. Moisture in breath reacts to make oxygen.

$$4KO_2(s) + 2H_2O(l) \rightarrow 3O_2(g) + 4KOH(aq)$$

A subsequent reaction removes exhaled carbon dioxide.

$$CO_2(g) + KOH(aq) \rightarrow KHCO_3(aq)$$

2. Sodium vapor lamps are used to illuminate roadways.
3. Oxides of calcium are used to neutralize acids in lakes affected by acid rain.

Extensions

1. Alkali metals have much higher than average conductivity both of heat and electricity. The high thermal conductivity of sodium, low melting point, and low price explain why sodium is often used as the heat exchange medium in nuclear reactors. How safe is it, given the high reactivity of elemental sodium?
2. Dietary requirements for sodium ions (Na^+) and potassium ions (K^+) can be located in many references. These ions can be found in many foods. The primary sources of K^+ are fruits and vegetables. Potassium ions aid the conduction of nerve impulses. What is the value of sodium ions in our diet. How much is healthy? How much is too much?
3. Students can survey labels on health foods to discover the levels of salt and monosodium glutamate in them. Lean Cuisine™, for example, has very high levels of sodium ion (Na^+) listed on its labels. What does "low salt" on a label mean?
4. NaCl is used on roads as a "de-icer." Consider the benefits and risks.
5. Research the Solvay Process. (See *Demonstration 3, Industrial Inorganic Chemistry* module, and Schematic for Industrial Preparation of Na_2CO_3 in the *Appendix* for more details.)
6. The student is given an unknown white solid. The goal is to identify the cation present in the salt. You might review the demonstrations (flame tests) seen in order to help develop a procedure to achieve this goal. The unknown might consist of Li, Na, K, or Cs salt(s). This project can be given as a paper and pencil exercise, or an actual laboratory research project.

References

Module developed by Herbert Bassow, William Bleam and Arthur C. Breyer, the Pennsylvania team.

Agricola, G. (1958). *De re metallica* (H. C. & L. H. Hoover. Trans.). New York, NY: Dover. (Original work published 1556).

This work covers mining and metallurgy with a section in Chapter 12 on salt manufactured from sea water and springs and wells.

Brooks, D. W. (Producer). (1989). *Doing chemistry* [Videodiscs, computer program and supporting written materials]. Washington, DC: American Chemical Society.

Brown, T., and LeMay, H. E. (1988). *Chemistry: The central science* (4th Ed.). Englewood-Cliffs, NJ: Prentice-Hall.

Properties (pp. 150–152) and reactivity (pp. 213–214) of alkali metals are discussed.

Chemical Education Material Study (Chem Study). (1969). *Chemistry, An experimental science*. New York, NY: Freeman.

Chapter 6, "Structure of the atom and chemical periodicity," provides useful information of periodic trends.

Cotton, F. A., Wilkinson, G., and Gaus, P. L. (1987). *Basic inorganic chemistry* (2nd Ed.). New York, NY: Wiley.

Chapter 10, "The Group IA (1) Elements" contains some interesting chemistry and excellent questions.

Gillespie, R., Humphreys, D., Baird, N. C., and Robinson, E. (1989). *Chemistry* (2nd Ed.). Boston, MA: Allyn & Bacon.

Chapter 15, "The Alkali & Alkaline Earth Metals" and a discussion of the photoelectric effect are pertinent to this module.

Greenwood, N. N., and Earnshaw, A. (1984). *Chemistry of the elements*. New York, NY: Pergamon Press.

"Lithium, Sodium, Potassium, Rubidium, Cesium and Francium" (Chapter 4) is a very detailed presentation on the alkali metals.

Haddad, N. and Tempest, R. (September, 1993). Simplified but reliable flame spectroscopy, *Chem 13 News*, p. 16.

Holtzclaw, H., and Robinson, W. (1988). *General chemistry* (8thEd.). Lexington, MA: Heath.

Chapter 13. "The Active Metals," provides good coverage of occurrence, preparation, reactions, and compounds; treats alkali metals, alkaline earth metals and aluminum group in an interesting and effective comparative way.

Kotz, J., and Purcell, K. (1987). *Chemistry and chemical reactivity*. Philadelphia, PA: Saunders.

Interesting sections of this text are pp. 756–764 on "The Alkali Metals: Group IA" and pp. 779–780, special section on fireworks.

Lagowski, J. (1973). *Modern inorganic chemistry*. New York, NY: Marcel Dekker.

Chapter 8, "The Alkali Metals," presents a more specialized chemistry of these elements.

McQuarrie, D., and Rock, P. (1985). *Descriptive chemistry*. New York, NY: Freeman.

"The Alkali Metals"(pp. 19–27) has good photos and data tables.

Multhauf, R. (1978). *Neptune's gift*. Baltimore, MD: Johns Hopkins University Press.

The history of common salt, table of salt production over the years by various countries, history and applications of salt, chemistry of salt are included.

Murray, P. R. S., and Dawson, P. R. (1976). *Structural and comparative inorganic chemistry*. London: Heineman Educational Books.

Chapter 8, "Group 1: The Alkali Metals" has some useful tables and graphs.

Steele, D. (1966). *The chemistry of the metallic elements*. New York, NY: Pergamon Press.

Chapter 4, "Group IA: The Alkali Metals," contains very interesting tables on chemical reactivity.

Weeks, M. E. (1968). *Discovery of the elements* (7th Ed.). Easton, PA: Journal of Chemical Education.

This book (the 7th edition was revised by H. Leicester) is now out of print. It would be a very worthwhile addition to your library if you could obtain it from a second-hand bookstore.

Whitten, K. W., Gailey, K. D., and Davis, R. E. (1988). *General chemistry with qualitative analysis* (3rd Ed.). Philadelphia, PA: Saunders.

Chapter 22, "Metals and Metallurgy" and Chapter 23, "The Representative Metals" are excellent.

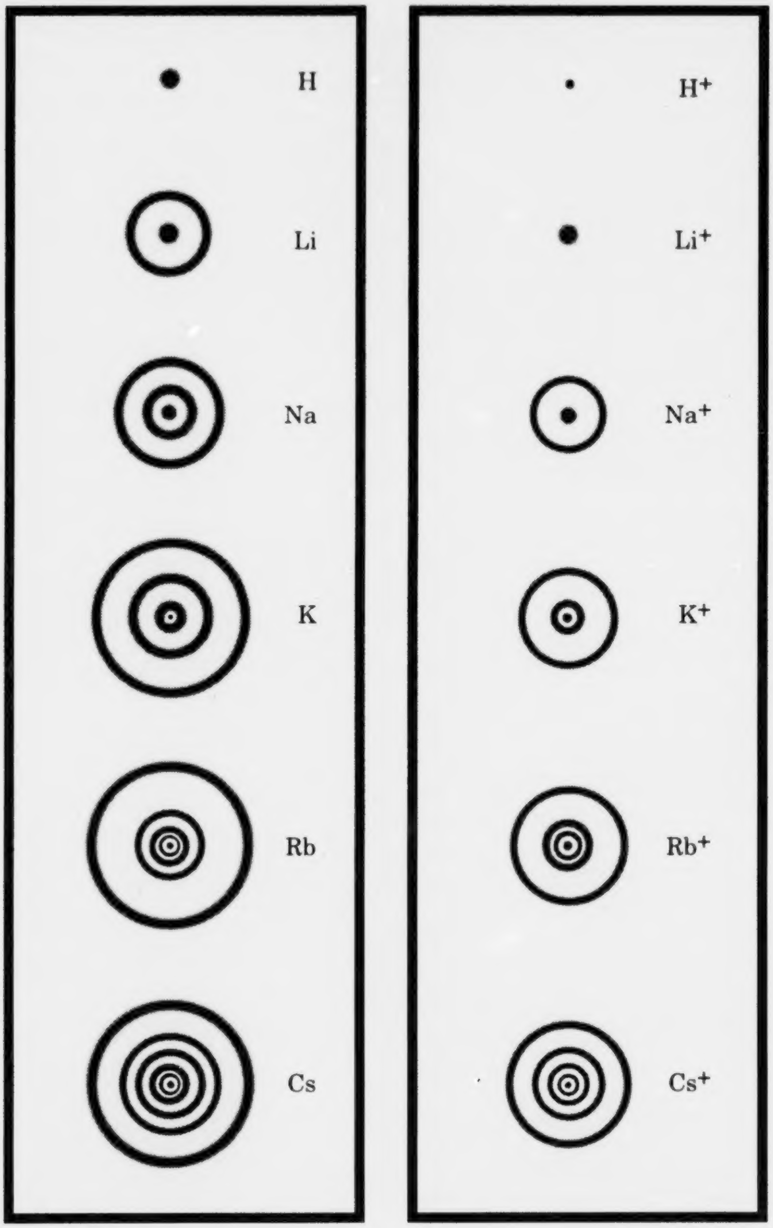
Zumdahl, S. S. (1986). *Chemistry*. Lexington, MA: Heath.

"The Properties of a Group: The Alkali Metals" (pp. 280–284) and the special fireworks section (pp. 285–287) provide interesting material.

Appendix

- **Transparency Masters**
  1. Cross Section of Alkali Metal Atoms and Ions
  2. Properties of the Alkali Metals
  3. Word Search
- **Humor**

Cross-Section of Alkali Metal Atoms and Ions


H
Li
Na
K
Rb
Cs
H+
Li+
Na+
K+
Rb+
Cs+

Selected Properties of the Alkali Metals

Element	Li	Na	K	Rb	Cs	Fr
Atomic Number, Z	3	11	19	37	55	87
Electronic Structure	$[He]2s^1$	$[Ne]3s^1$	$[Ar]4s^1$	$[Kr]5s^1$	$[Xe]6s^1$	$[Rn]7s^1$
Atomic Radius, pm	123.0	157.0	203.0	216.0	235.0	
Atomic Mass, u	6.941	22.9898	39.0983	85.4678	132.9054	(223)
1st Ionization Energy, kJ/mol	520.1	495.7	418.6	402.9	375.6	~(375)
2nd Ionization Energy, kJ/mol	8067.0	5844.0	3393.0	2970.0	2692.0	
Ionic Radius (Pauling), pm	60.0	95.0	133.0	148.0	169.0	
Electron Affinity, kJ/mol	–54.8	–52.7	–48.36	–46.89	–45.50	
Electronegativity (Pauling)	1.0	0.9	0.8	0.8	0.7	0.7
Covalent Radius, pm	128.0	156.0	198.0	214.0	230.0	
Density, g/cm^3	0.534	0.968	0.85	1.532	1.90	
Melting Point, °C	186.5	97.5	62.3	38.5	28.5	27.0
Boiling Point, °C	1336.0	881.4	765.5	688.0	705.0	677.0
Heat of Fusion,kJ/mol	2.93	2.64	2.39	2.20	2.09	
Heat of Vaporization, kJ/mol	148.0	99.0	79.0	76.0	67.0	
Heat of Atomization, kJ/mol	161.0	109.0	90.0	86.0	79.0	
MOH hardness	0.6	0.4	0.5	0.3	0.02	

Selected Properties of the Alkali Metals (continued)

Element	Li	Na	K	Rb	Cs	Fr
Ionic Potential Charge (+1)/Ion Radius, nm^{-1}	16.7	10.5	7.5	6.8	5.9	
$\frac{\text{Charge Density}}{\text{Ion Surface Area}}$ (for +1 charge)	22.0	8.8	5.9	3.6	2.8	
(Pauling's values) Hydrated Ionic Radius, pm	340.0	276.0	232.0	228.0	228.0	
Hydration Number (H_2O's per M^+ ion)	25.3	16.6	10.5	---	9.9	
Equivalent Conductivity x 10^{-4} m^2 Siemens mol^{-1} at 18 °C	33.5	43.5	64.6	67.5	68.0	
Equivalent Conductivity x 10^{-4} m^2 Siemens mol^{-1} at 25 °C	38.6	50.2	73.5	77.8	77.3	
Reduction Potential (E°), V	–3.03	–2.713	–2.92	–2.93	–2.92	
Crystal Structure	BCC	BCC	BCC	BCC	BCC	
Metal Radius, pm	153.0	188.0	231.0	248.0	266.0	~290
Critical Temperature, K	3233.0	2509.0	2220.0	2083.0	2051.0	
Critical Pressure, MPa	68.9	25.6	16.4	14.5	11.7	
Heat of Dissociation of $M_2(g)$, kJ/mol	107.8	73.3	49.9	47.3	43.6	
Thermal Conductance, J/s cm^2 °C/cm^3	0.71	1.326	0.97			
Entropy,J/mol K	28.0	51.5	63.6	69.5	82.8	
Flame Color	crimson	yellow	violet	red-violet	blue	
Persistent Wavelength, nm	670.8	589.2	766.5	780.0	455.5	
Abundance in Sea water, *M*	6×10^{-5}	0.47	0.01	1×10^{-1}	$\sim 1 \times 10^{-8}$	

Selected Properties of the Alkali Metals (continued)

Element	Li	Na	K	Rb	Cs	Fr
Ionic Radius, pm (coordination number = 6)	76.0	102.0	138.0	152.0	167.0	(180)
Effective Number of free electrons per atom	0.55	1.1	0.97	0.94	0.85	
Covalent Radius, pm	128.0	156.0	198.0	214.0	230.0	

Thermal Conductivity (Watt $cm^{-1}K^{-1}$)			
	Temperature (K)		
	273.2	298.2	373.2
Li	0.859	0.848	0.818
Na	1.42	1.42	1.32
K	1.036	1.025	—
Rb	0.583	0.582	0.581
Cs	0.361	0.359	—

Specific Electric Conductivity (cm^{-1} ohm^{-1})		
	Temperature (K)	
	273.2	298.2
Li	1.17×10^5	1.08×10^5
Na	2.31×10^5	2.10×10^5
K	1.54×10^5	1.39×10^5
Rb	8.67×10^4	7.79×10^4
Cs	5.33×10^4	4.89×10^4

Specific Resistivity (ohm m)		
	Temperature (K)	
	273.2	298.15
Li	8.53×10^{-8}	9.28×10^{-8}
Na	4.33×10^{-8}	4.77×10^{-8}
K	6.49×10^{-8}	7.20×10^{-8}
Rb	11.54×10^{-8}	12.84×10^{-8}
Cs	18.75×10^{-8}	20.46×10^{-8}

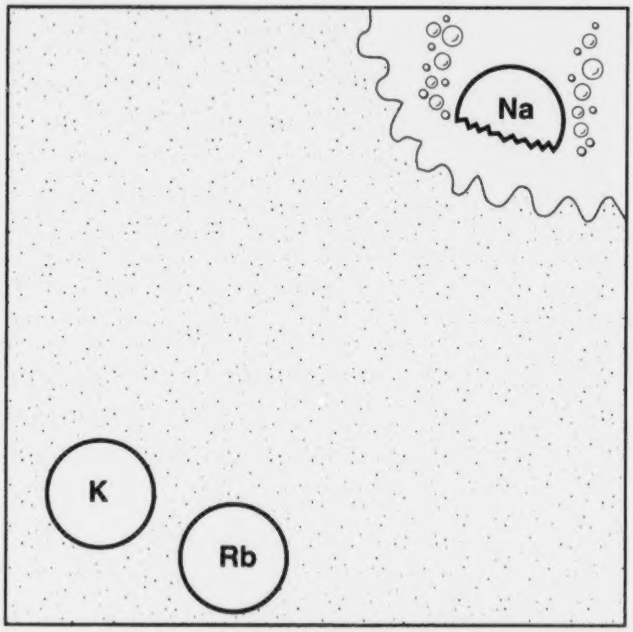
Reference: *CRC Handbook of Chemistry and Physics, 7th Ed.* (1989-90). pp. 146-147.

Word Search

E C Y D T W A O R B P N P K A Q V
T X F T B G N U E D E C G U T Z N
S P B R I Q I H M V E M V U C H D
U F I Z X S O T D S L Q C N R W N
I N Q G O D N U N I T C E L L P H
D I C A A O R E R H P F A F H T N
A O P N I O G S D Y K Q R D N C W
R M Y T E O U X N E L K M T L Q A
C I A J L L G L H B G S E L W F J
I C Q A F S I S Y L O R T C E L E
N N H C S I Y A G L K V A S O V X
O K Y S G O T H Y M N U L H N Y J
I A L K A L I M E T A L S Q C D I

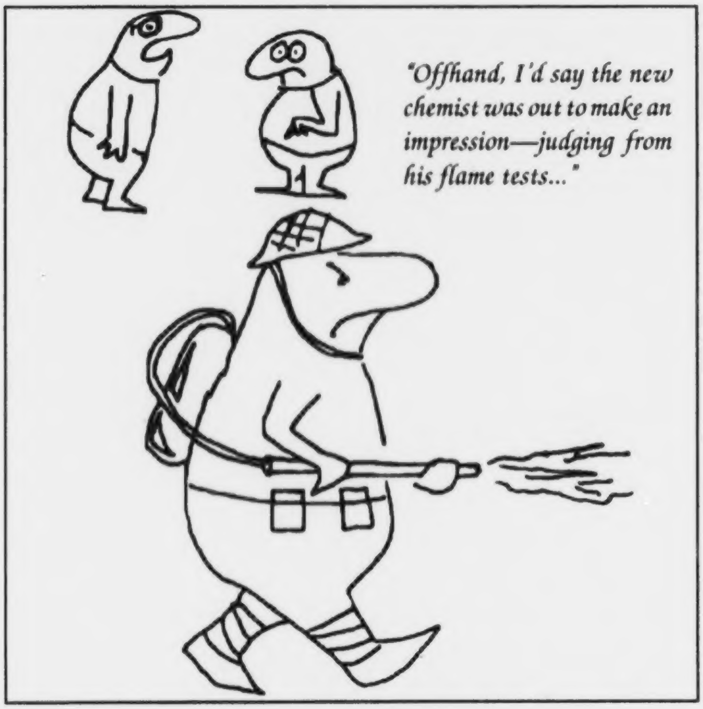
Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Family of elements characterized by its vigorous reaction with water (2 words).
2. Negatively charged ion.
3. Positively charged ion.
4. Charge on an ion divided by its surface area (2 words).
5. Process by which compounds are decomposed electrically.
6. Family that is very reactive with alkali metals in which binary ionic compounds in a one-to-one ratio are formed.
7. Property of alkali metals that increases from Li^+ to Fr^+ (2 words).
8. Group of elements characterized by luster, malleability, ductility, and good electrical and heat conductivity.
9. Mineral deposit containing sufficiently high concentration to allow practical recovery of a metal.
10. Smallest part of a crystal that, if repeated, could generate the whole crystal (2 words).

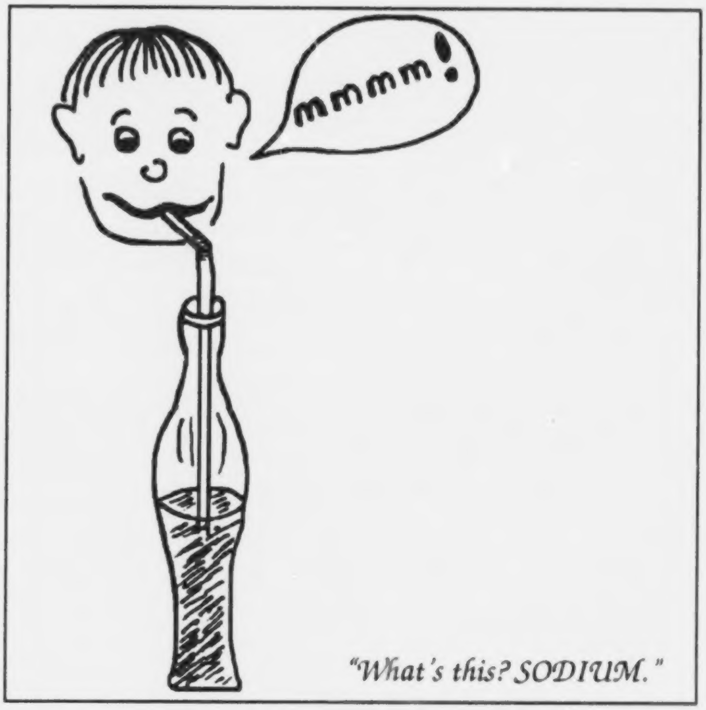


I told him, "Nat, don't go near the water."
But nobody ever listens to me.

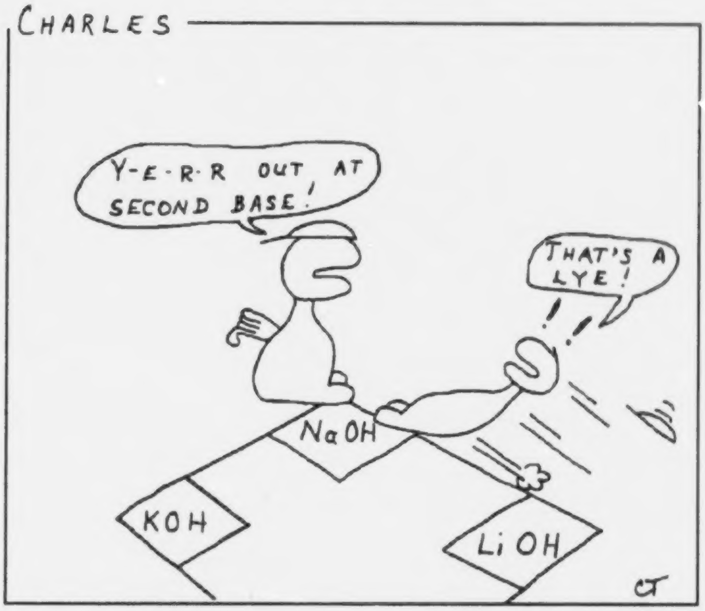
© John Ihde, Wausau West HS, Wausau, WI 54401. Used by permission.


"Offhand, I'd say the new chemist was out to make an impression—judging from his flame tests..."

CHEM 13 NEWS, April 1978, p. 1351. Reprinted with permission.

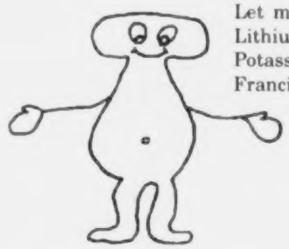

mmmm!
"What's this? SODIUM."

CHEM 13 NEWS, March 1983, p. 9. Reprinted with permission.

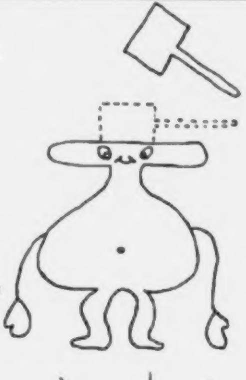

CHARLES
Y-E-R-R OUT, AT SECOND BASE!
THAT'S A LYE!
!!
NaOH
KOH
LiOH
CT

CHEM 13 NEWS, October 1983, p. 16. Reprinted with permission.

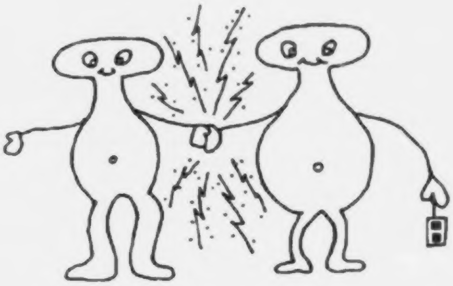
The Tale of the Alkali Metals



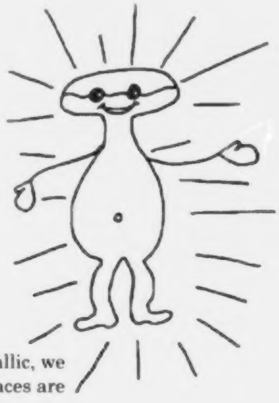
Let me tell you a little about me, Lithium, and my brothers Sodium, Potassium, Rubidium, Cesium, and Francium.



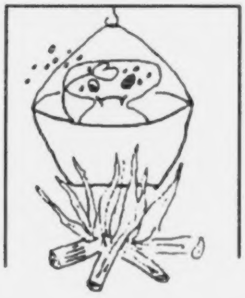
All of us alkali metals are malleable (can be re-shaped with pressure).



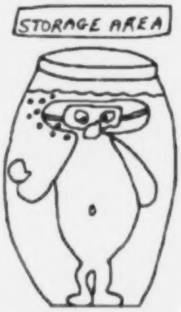
All of my family make good electrical conductors.



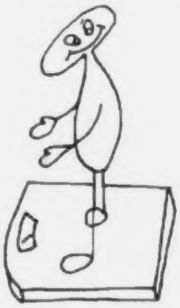
Because we are metallic, we shine when our surfaces are clean.



My family has the lowest melting points out of all our other relatives that live on the Periodic Table.



All of my family, except for me, Lithium, have to live under kerosene or some other similar liquid.



My family are quite light in weight (low density).

CHEM 13 NEWS, February 1977, p. 10-12. Reprinted with permission.

A SourceBook Module

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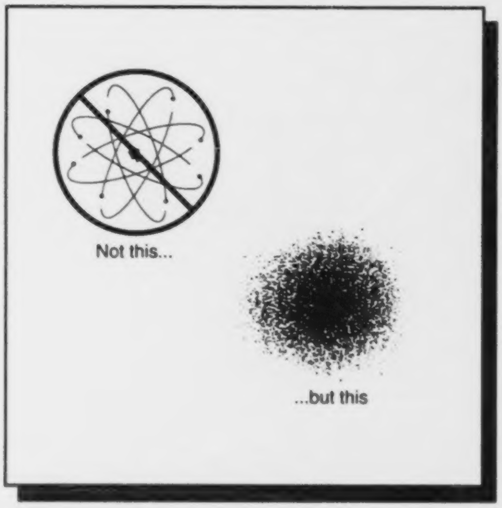
ATOMIC STRUCTURE



ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Atoms are composed of three smaller parts—protons, neutrons, and electrons. Protons are positively charged, neutrons have no charge, and electrons are negatively charged. A neutral atom contains an equal number of protons and electrons. The protons and neutrons together form the nucleus of the atom, occupying less than 0.01% of the total atomic volume but accounting for 99.99% of the atomic mass. Electrons, although contributing negligibly to the mass, spend their time in the remaining volume. This means that most of the volume of an atom is essentially "nothingness." The number of protons in the nucleus of an atom determines the element, but the electrons determine the element's chemical behavior. The outermost or valence electrons are those that can be shared with other atoms to form molecules, or can be transferred to or from other atoms, forming ionic salts. To understand how valence electrons determine the chemistry of an element, the internal structure of the atom must be examined in greater detail.

Electrons occupy specific "spaces" or energy levels in an atom. These energy levels are quantized; the location of electrons in terms of energies (the electron configuration) determines the nature and number of chemical bonds that each element forms. The chemical and physical properties of an element are determined by its electron configuration. These properties, the direct consequence of electron configuration, are observed in such phenomena as fireworks, neon lights, television, and colorful fireplace logs. In each instance, the distinctive colors are due to excitation of valence electrons.

PLACE IN THE CURRICULUM

This topic can be introduced early in the course and should precede modules on *Periodicity*, *Oxidation-Reduction*, and *Bonding*. It requires the use of fairly abstract concepts that may need additional reinforcement periodically in the course.

CENTRAL CONCEPTS

1. An atom is composed of a nucleus containing protons and neutrons, which is surrounded by electrons. The electrons, though of little relative mass, occupy most of the atomic volume.
2. Each electron and proton possesses the same absolute electrical charge value, though of opposite sign. By convention, electrons are assigned a negative charge (−) while protons are assigned a positive charge (+).
3. The *atomic number* represents the number of protons in the nucleus of an atom. The *mass number* represents the sum of protons and neutrons in the nucleus. The *atomic mass* is the weighted average of all naturally occurring isotopes (same atomic number, different number of neutrons) of an element.
4. Light has a dual nature, behaving both as waves and particles. This behavior is governed by several mathematical relationships:

$$c = \lambda \nu$$
$$E = h \nu$$

c = Speed of light
λ = Wavelength of light
ν = Frequency of light
E = Energy per photon of light
h = Planck's constant

5. Electrons have a dual nature, behaving both as waves and particles.
6. Electrons occupy specific energy levels in an atom. These energy levels are quantized and are described by shells, subshells, and orbitals.
7. The orbital occupancy of a multi-electron atom is given by its electron configuration. This electron configuration can be described as following several empirical rules—Hund's rule, the Pauli exclusion principle, and the aufbau principle.
8. The structure of the Periodic Table is a reflection of the electronic structure of atoms.

RELATED CONCEPTS

Students should have previously studied the particulate nature of matter and the concept of atoms as "building blocks" of matter.

RELATED SKILLS

Use of calculator for exponential numbers.

PERFORMANCE OBJECTIVES

After completing their study of atomic structure, students should be able to:

1. show the interrelationships of atomic number, atomic mass, numbers of protons, neutrons, and electrons.
2. describe the relative energies of ultraviolet, visible, infrared, microwave, X-ray, radio, and TV waves.
3. determine the electron configuration of a specified neutral atom.
4. draw an orbital diagram that correlates with a given electron configuration.
5. identify the maximum number of electrons that can occupy a shell or set of orbitals.
6. distinguish between absorption (excitation) and emission of energy.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Identification of Substances through Atomic Emission of Light

(Adapted from Laboratory Experiments for Foundations of Chemistry by Ernest R. Toon and George L. Ellis, Holt, Rinehart, and Winston (1982), pp. 106-109—no longer in print.)

Introduction

Light is emitted by gases or vapors when excited. Light and other kinds of electromagnetic radiation are forms of energy. When energy is converted to radiation and light is given off, emission has occurred. When an atom or molecule gains or loses energy, this energy is transferred to and from atoms or molecules in definite, discrete units of radiant energy called photons or quanta. The energies of electrons in an atom or molecule are restricted to definite quantities, often called energy levels. These energy levels are fixed for a given atom. To change the energy state, a photon must be emitted or absorbed. In other words, electrons move from one energy level to another.

In this activity you will discover that the emission spectrum of an element is characteristic of that element. The lines represent quantities of energy emitted when excited electrons move from one energy level to another. Sodium, as an example, has two closely spaced visible lines.

Purpose

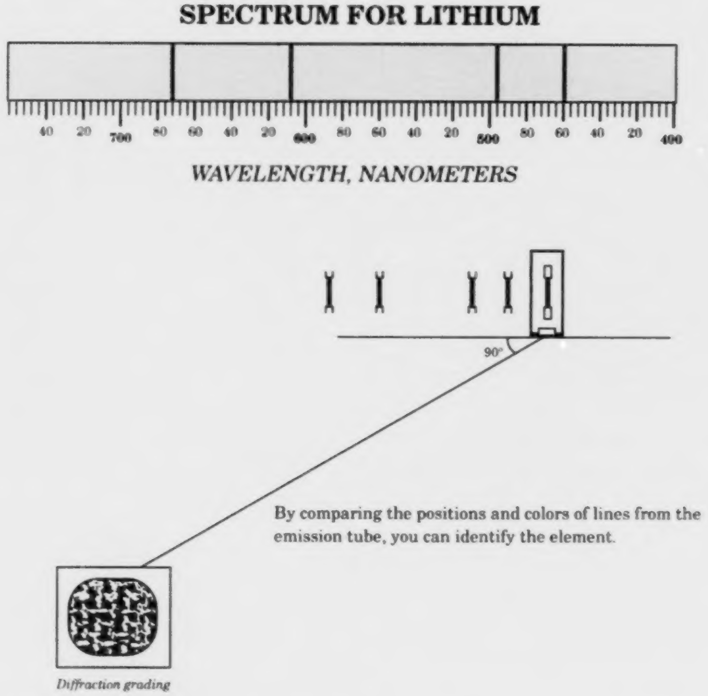
In this laboratory activity you will identify the spectra of unknown elements and calculate the frequency and energy of particular lines found in the spectra.

Safety

1. Electrical shock is possible if the metal clamps are touched while the power supply is on.
2. The gas tubes must be handled very carefully.
3. Excited elements emit ultraviolet light, which can harm your eyes. Avoid looking directly at the spectrum tubes!

Procedure

1. Set up the apparatus shown in Figure 1.
2. Observe the spectral lines from a known-element spectrum tube. The spectral lines can be viewed by holding a diffraction grating to your eye and looking at the spectrum tube. Spectral lines will appear on both sides of the spectrum tube. Observe only those lines closest to the spectrum tube (first-order lines).
3. Compare the colors and relative distance of the lines from the spectrum tube to lines found on the spectrum chart. Confirm that the observed spectrum matches the same colors and patterns as that shown on the spectrum chart.
4. Turn off the power supply. Remove the "known" spectrum tube and insert the "unknown" spectrum tube.
5. By noting the color and pattern of the emission lines from the spectrum tube, determine which element the unknown tube is and the wavelength of at least three of the spectrum lines. Record your results in a data table.


SPECTRUM FOR LITHIUM
40 20 700 80 60 40 20 600 80 60 40 20 500 80 60 40 20 400
WAVELENGTH, NANOMETERS
90°
Diffraction grading

By comparing the positions and colors of lines from the emission tube, you can identify the element.

Figure 1. Spectrum for Lithium.

Data Analysis and Concept Development

Using the following mathematical relationships, answer Questions 1 and 2.

$E = h\nu$
$c = \lambda\nu$

1. Calculate the frequency (ν) of a light source with a wavelength of 600 nm.
2. Calculate the energy (E) from the frequency calculated in Question 1. ($h = 6.63 \times 10^{-34}$ J s)

Implications and Applications

1. What causes spectral lines?
2. What is meant by the term monochromatic?
3. How would you compare the accuracy of this method of determining the identity of an element to that of a flame test?
4. Why do elements have several spectral lines?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Identification of Substances through Atomic Emission of Light

Major Chemical Concept

Characteristic light emissions can provide evidence regarding the identity of many elements.

Level

This activity can be completed by students at any level. Post-laboratory calculations should not be too difficult for the basic student if they are given proper preparation.

Expected Student Background

Students should review atomic structure and study a spectrum chart of the elements. Discuss the different types of spectroscopes and their uses in chemical analysis. This activity can follow *Demonstration 7* on flame tests and be presented during the discussion of atomic structure and quantum number.

Time

If students do three unknowns, they can be completed in one laboratory period.

Safety

See *Student Version* safety pointers. When handling these tubes, be sure to disconnect the electricity. Under no circumstances allow the students to handle the tubes. Electrocution can result!

Materials (For 24 students working in pairs)

Diffraction grating (Frey Scientific, Catalog #188, 10 mounted gratings for $5.95, 1-800-255-FREY); alternatively, a holographic diffraction grating sheet can be purchased from Learning Technologies, Inc., 59 Walden Street, Cambridge, MA 02140; (800) 537-8703; (617) 547-7724. One-cm square sections can be mounted over a small hole punched in an index card. Very economical!
Spectrum tubes (Na, He, H_2, Hg); use one as the "known" (Frey Scientific or Fisher)
Power supply for the spectrum tubes
Visible Range Spectrum Chart (Fisher Scientific, Catalog No. 05-712-50)

Advance Preparation

Inspect the spectrum tubes to confirm that they emit enough light. If not, you may have to darken the room.

Pre-Laboratory Discussion

Discuss the different types of spectroscopes and their uses in chemical analysis. Demonstrate how diffraction gratings work. Pass gratings among students and have them look at lights in the room. You may wish to discuss how prisms and gratings disperse light. Standard physics textbooks carry explanations of these two different phenomena.

Teacher-Student Interaction

1. Check to make sure the student is using the grating properly. Can they see the lines? If not, try having them rotate the grating 90° and look to the side. Possibly the tube is old and a weak emitter, or the student is not close enough to it (less than 10 feet is best).
2. Make sure students do not smudge the gratings with fingerprints.

Answers to Data Analysis and Concept Development

1. Given a wavelength (λ) of 600 nm calculate the frequency (ν).

   a. Change the wavelength, λ, from nm to m: $600 \text{ nm} \times \frac{1\text{m}}{10^{-9}\text{ nm}} = 6.00 \times 10^{-7}\text{ m}$

   b. Calculate frequency (ν) $\nu = \frac{c}{\lambda} = \frac{3.00 \times 10^{8}\text{ m/s}}{6.00 \times 10^{-7}\text{ m}} = 5.00 \times 10^{14}\text{ s}^{-1}$

2. $E = h\nu$

   $= (6.63 \times 10^{-34}\text{ J·s})(5.00 \times 10^{14}\text{ s}^{-1})$

   $= 3.32 \times 10^{-19}\text{ J}$

Answers to Implications and Applications

1. Visible spectra are produced by electron transitions. When an element such as sodium is heated, the valence electron moves from one energy level to another. The electron absorbs energy when it moves from a lower to higher energy level and emits light when it drops to a lower energy level.
2. Monochromatic means one color; the light is of one wavelength. Laser light is monochromatic.
3. A flame test may give a variety of colors. Students may have difficulty making correct color observations. However, spectral lines are known values that can be used for a more positive means of identification. Spectral lines are quantitative; the wavelength, frequency and energy of lines can be calculated from experimental data with great accuracy.
4. Any given element will have a number of energy levels. Therefore, as electrons move to and from different energy levels, a series of spectral lines will be produced.

Assessing Laboratory Learning

1. What causes spectral lines? *[See answer to Question 1 from* Implications and Applications.*]*
2. Why are the spectra of two elements different? *[The different numbers of protons in the nucleus and electrons in the orbitals leads to different energy differences between levels in each element. A spectrum is like a fingerprint for an element.]*
3. A sodium vapor lamp, similar to many yellow streetlights, emits light of wavelength 589 nm. What are the energy and frequency of this light? *[Answer:* $E = 3.37 \times 10^{-19}$ *J,* $\nu = 5.10 \times 10^{14}\text{ s}^{-1}$. *Calculations are similar to 1 and 2 in* Data Analysis.*]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Locating an"s" Electron in an Atom by Analogy

Introduction

If it were experimentally possible to determine how often an electron appeared at different points in space surrounding the nucleus, these data could be plotted on a three-dimensional graph. Boundaries could be established that would outline the regions of space within which the electron could be found 95% of the time. These regions of space, called orbitals, are commonly represented in texts as charge clouds, and physically through Styrofoam models.

Since it is impossible to conduct such an experiment, we shall perform one that, by analogy, will help give meaning to the concepts of probability distribution graphs, radii of maximum probability, maximum electron density, and orbitals.

In this laboratory activity, you will define regions of space around a bulls-eye in which there is a specific probability of locating a spot resulting from the impact of a marble dropped from a specified distance to the target. The spots represent points in space around the bulls-eye (analogous to the nucleus) where the marble (analogous to the electron) is observed in an experiment to locate the region of space within which the marble is most likely to strike (analogous to an atomic orbital).

You will also determine a radius of maximum probability (maximum spot density) by plotting the number of times the spots (analogous to places electrons might be observed) appear times the radius *vs.* distance from the bulls-eye (nucleus).

Purpose

1. To draw a physical analogy to some concepts associated with the wave-mechanical model of the atom.
2. To determine the distribution of impacts of marble drops around the bulls-eye of a target.
3. To obtain and interpret probability information on the distribution of marble drops around a bulls-eye of a target.

Procedure

1. Obtain two targets, a piece of typewriter carbon paper, and four paper clips.
2. Place the carbon paper between the two targets so that any impression made on the front target will be reproduced on the back target. Fasten with a paper clip at each corner.
3. With care, stand on a stool or laboratory bench with the marble in your hand. Extend your arm so it is about two meters above the floor. Have your partner place the target directly below your hand. Drop the marble on the target. Your partner must catch the marble on the *first* bounce (otherwise you'll have more than one mark for each drop of the marble). Repeat 99 times. A marble drop location will thus appear 100 times. The spots on the back target represent places where the marble was observed.
4. Return the front target, carbon paper, marble, and paper clips to locations designated by your teacher.

Data Analysis and Concept Development

1. Locate the 95 dots (95% of the 100) that are closest to the bulls-eye. Using a pencil, draw a smooth (not wavy) curve to enclose the region containing these dots. This region represents the two-dimensional "orbital" of your "marble electron" (analogous to the region of space where an electron might be observed 95% of the time in an experiment to locate it in three-dimensional space).
2. Count the dots in each concentric circular area surrounding the bulls-eye and record the number in the table. Multiply the number of dots in each area by the radius of the region given in the table. The product represents the dots times centimeters. Enter your calculations in the table.

Number of circular area increment (n)	Maximum radial distance of area increment from bullseye	Area of increment*	Number of dots in radial increment	Dots times area
0	0	0	0	0
1	R_1 = 0.9 cm	2.54 cm^2		
2	R_2 = 2.6 cm	18.7 cm^2		
3	R_3 = 4.4 cm	42.1 cm^2		
4	R_4 = 6.3 cm	82.6 cm^2		
5	R_5 = 8.0 cm	119 cm^2		

*Area of increment = $\pi R^2_{n+1} - \pi R^2_n$

Radius of maximum probability (from graph) ___________

Figure 2. Score card.

3. Prepare a graph by plotting the dots times centimeters along the vertical axis and the maximum radial distance of the area increment from the bulls-eye along the horizontal axis. Note that radial distance extends from the center of the bulls-eye to the midpoint of the circular area. Start the graph line at 0,0 and allow it to approach the *x* axis asymptotically (without touching the axis).
4. Identify the radius of maximum probability from the graph (peak of the curve) and record it in the space provided in the table. This is the distance from the bulls-eye where you are most likely to find a spot (marble or electron) when attempting to locate it (drop a marble).
5. Compare your results to the radial probability curve for the *s*-electron (Figure 3). Sketch the *s*-electron plot on the graph paper containing your experimental curve. Use a colored pencil for the *s*-electron curve.

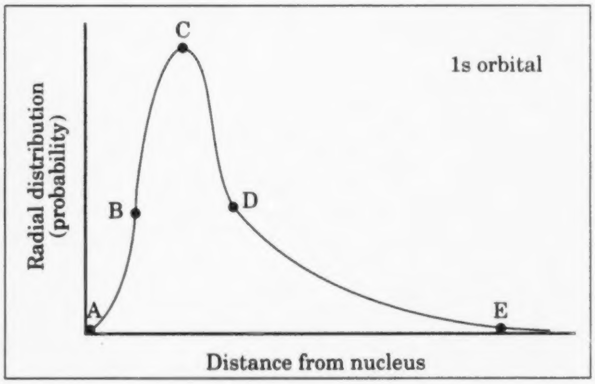


Figure 3. Radial probability curve for the s electron.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Locating an "s" Electron in an Atom by Analogy

Major Chemical Concept

This laboratory exercise involves no "chemistry". It is designed to give students insight into the meaning of abstract concepts and terms (orbitals, electron density, charge cloud, probability graphs, *etc.*) associated with the wave-mechanical model of an atom. This activity is a simulation of the radial distribution pattern that helps calculate the Bohr radius. However, it does not represent reality because the radial distribution pattern is actually obtained by a calculation involving the differential volume of a sphere.

Level

The activity can be done by students at all levels; the abstract concepts can be discussed at a variety of levels, depending on the abilities of the students.

Expected Student Background

Students should review wave-mechanical model of the atom.

Time

One laboratory period.

Safety

There are no unusual safety requirements, except to insure that students with marble-drop responsibilities climb carefully onto stable, safe structures.

Materials (For 24 students working in pairs)

24 Targets (see sample targets in *Appendix*)
12 Typewriter carbon papers
48 Paper clips
12 Marbles

Advance Preparation

Obtain enough marbles, targets, and carbon paper.

Pre-Laboratory Discussion

The materials and procedures of the activity may be illustrated during a brief pre-laboratory discussion. During this discussion, explain that the activity's purpose is to develop "experimentally" a "charge cloud" diagram and probability graph. Point out the analogy to comparable mathematically derived electric charge-cloud diagrams and probability graphs found in some student texts. *NOTE: Student graphs should begin at a radius of 0, where the probability of finding an electron is 0.*

Teacher-Student Interaction

1. Be sure the marbles are caught after their first bounce.
2. Be sure marbles are dropped from a consistent height (2 m).
3. Ask students whether they expect each group to get the same pattern (the patterns should be similar, but not identical).
4. Do students expect a marble to land *exactly* in the middle?

Answers to Questions in Data Analysis and Concept Development

Number of circular area increment (n)	Maximum radial distance of area increment from bullseye	Area of increment*	Number of dots in radial increment	Dots times area
0	0	0	0	0
1	R_1 = 0.9 cm	2.54 cm^2	1	2.54
2	R_2 = 2.6 cm	18.7 cm^2	39	729
3	R_3 = 4.4 cm	42.1 cm^2	36	1516
4	R_4 = 6.3 cm	82.6 cm^2	19	1569
5	R_5 = 8.0 cm	119 cm^2	5	595

*Area of increment = $\pi R_{n+1}^2 - \pi R_n^2$

Radius of maximum probability (from graph) ________

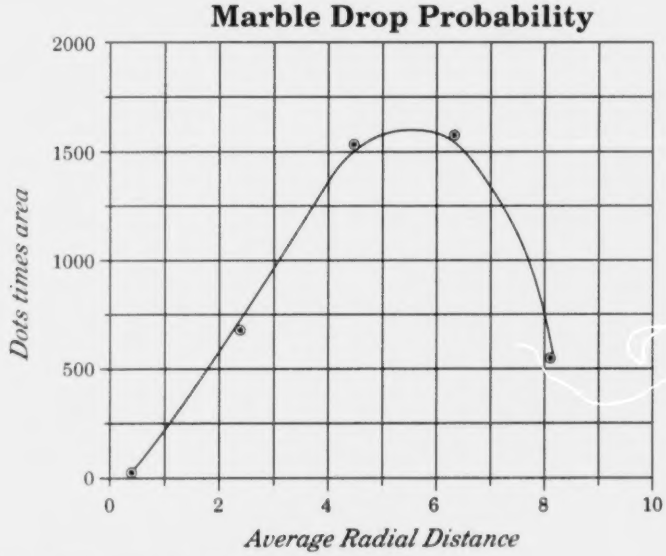


Figure 4. Sample data for Laboratory Activity 2.

Assessing Laboratory Learning

1. Why don't all the marbles land in the same spot? *[There are many uncertainties that change the landing spot each time, height of hand, angle of release,* etc.*]*
2. a. Which point (A-E) is the greatest probability of finding the electron? *[C]*
   b. Where is there zero probability of finding the electron? *[A: the nucleus]*
   c. As the distance from the nucleus increases, what happens to the probability of finding an electron? *[It peaks at some distance, C, then diminishes, but never reaches zero.]*

Obscertainer Kit from Laboratory-Aids

A simple laboratory activity to help students learn to infer patterns is available from Laboratory-Aids Inc., 249 Trade Zone Drive, Ronkonkoma, NY 11779 (phone 516-737-1133) for approximately $45 plus shipping (see sample worksheet in Appendix).

DEMONSTRATIONS

Demonstration 1: Spectra in the Real World

Give each student a small square of plastic diffraction grating to view H_2, He, Ne, *etc.* emission from spectrum tubes. Let students take the grating home to view sodium street lamps and mercury vapor lamps. *WARNING:* **Do not** *look at the sun!*

Demonstration 2: Getting a Charge out of Things

If students haven't seen the demonstration of an amber or glass rod (or a length of PVC pipe works very well) rubbed with fur, silk, or hair, and then used to attract or repel pith balls or other objects (such as suspended balloons), this is a good place to perform it.

Demonstration 3: Standing Waves

An old-style man's electric razor, with the cutting top removed, has an exposed vibrating post. By attaching a string to this post, you can show quantized standing waves when the razor is turned on.

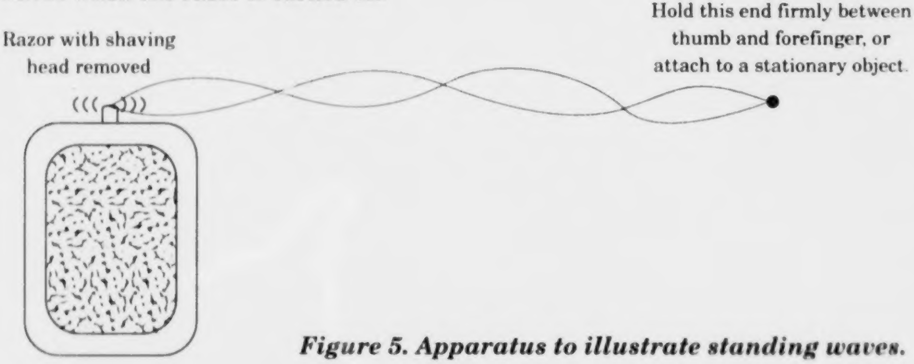

Razor with shaving head removed
Hold this end firmly between thumb and forefinger, or attach to a stationary object.

Figure 5. Apparatus to illustrate standing waves.

Demonstration 4: Modelling Orbitals

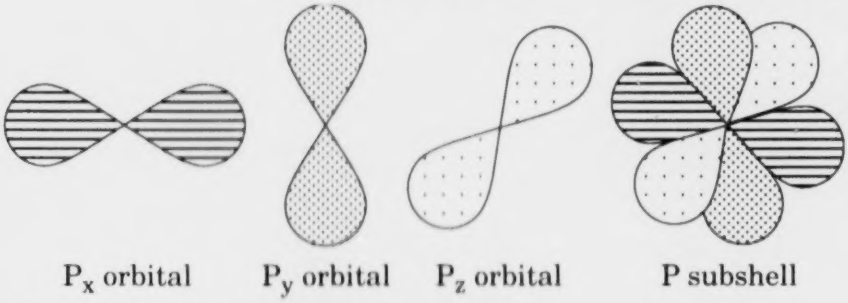


Figure 6. Balloons illustrating p-orbitals.

Spherical balloons can be used to model s-orbitals. Two spherical or oblong balloons tied together can be used to model p-orbitals.

Materials

6 Balloons (2 each of 3 different colors). Tie pairs of balloons (2 the same color) together. Bunch.

Demonstration 5: Wavelength and Color

To demonstrate the relationship between wavelength and color in the visible range from red to violet, use a spectrophotometer. Place a piece of white chalk in a spectrophotometer cell (see Figure 7). Cut (or sandpaper) the chalk at an angle that will reflect light upward to the student's eye looking down into the cell holder. Open the slit width. Rotate the wavelength knob to compare colors and wavelengths.

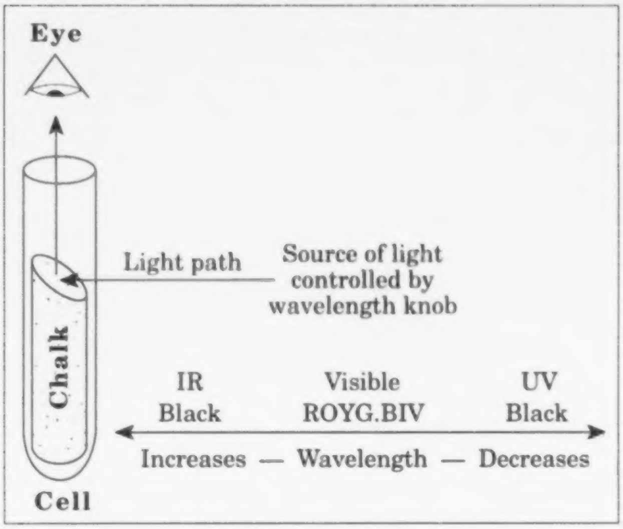

Eye
Chalk
Cell
Light path
Source of light controlled by wavelength knob
IR
Black
Visible
ROYG.BIV
UV
Black
Increases — Wavelength — Decreases

Figure 7. Demonstrating the relationship between color and wavelength.

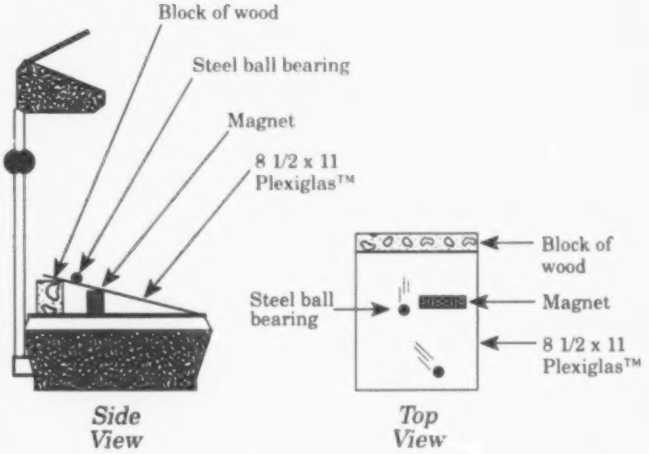
Demonstration 6: Simple Mass Spectroscopy

Materials

Overhead projector
Sheet of Plexiglas™, 8-1/2 x 11 in
Magnet, disk shape, 3-4 in diameter
3-4 Steel ball bearings of various sizes

Procedure

1. Form a shallow inclined plane with the Plexiglas™ sheet on top of the overhead projector, by using a small block of wood or other material to elevate one end of the sheet.
2. Insert the magnet underneath the center of the inclined plane.
3. Roll the steel balls down the incline. If they are rolled near but not over the magnet (they will stop) their direction of motion will be altered. The smaller ball bearings will be deflected through a greater angle than the larger ones.


Block of wood
Steel ball bearing
Magnet
8 1/2 x 11 Plexiglas™
Block of wood
Steel ball bearing
Magnet
8 1/2 x 11 Plexiglas™

Side View

Top View

Figure 8. Simple mass spectroscopy demonstration.

Demonstration 7: Flame Tests

Absorption and emission spectra are used to identify many elements. When elements are heated to high temperatures they may be placed in an excited state; in this excited state valence electrons move to higher energy levels. When the electrons return to their ground state, they may emit visible light of characteristic colors that can be used to identify the element.

In this demonstration, students will identify the colors of the emission spectra of some metallic ions, using a burner and wood splints.

Materials

Burner
Wood splints
Copper(II) nitrate, $Cu(NO_3)_2$, solid crystals
Barium nitrate $Ba(NO_3)_2$, solid crystals
Sodium nitrate, $NaNO_3$, solid crystals
Strontium nitrate, $Sr(NO_3)_2$, solid crystals

Procedure

Place a small amount of one of the solid ionic compounds (about the size of a rice grain) on the tip of a wooden splint. Place the splint at the edge of the hottest part of the flame (top of inner blue cone) and observe the color. (Some students, particularly some males, may be color blind. Consider allowing students to compare their observations in small groups, alerting them to the possibility of color blindness.)

(Note: Good results have been obtained with concentrated "syrups" of the nitrate salts.)

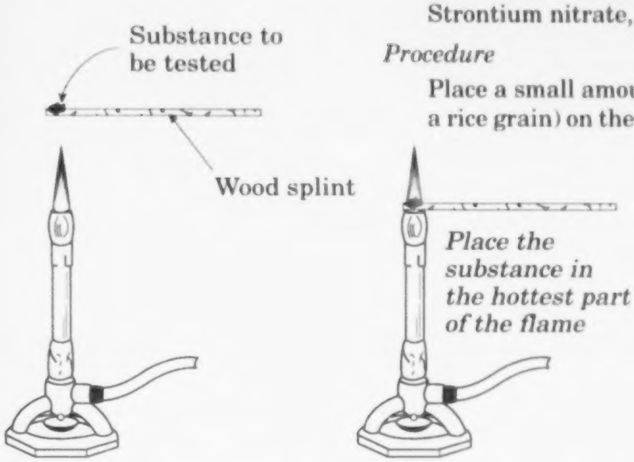


Figure 9. Flame test technique.

Results

Ionic salt	Metal ion	Color
Copper(II) nitrate, $Cu(NO_3)_2$	Cu^{2+}	Blue-green
Barium nitrate, $Ba(NO_3)_2$	Ba^{2+}	Green
Sodium nitrate, $NaNO_3$	Na^+	Yellow-orange
Strontium nitrate, $Sr(NO_3)_2$	Sr^{2+}	Bright red

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What are the principal parts of an atom? *[The nucleus and electron cloud. The nucleus contains protons, neutrons and many other subnuclear particles (see Nuclear Chemistry module). The electrons can be divided into core and valence electrons.]*
2. What forces hold atoms together? *[Coulombic (electrostatic) attractions between oppositely charged protons and electrons. Coulomb's law is* $F = k\frac{q_1q_2}{d^2}$*]*
3. How are electron "orbitals" different from "orbits"?
   [Orbitals refer to the probable location or region of space (a volume) where an electron is likely to be found with a specified degree of certainty. It arises from application of wave and uncertainty considerations. Orbits are Niels Bohr's representation of rigid circular or elliptical paths followed by particulate electrons.]
4. What is an energy level diagram? *[A graphic representation of the relative energies possessed by electrons in various orbitals as described by their quantum numbers.]*

5. How are electron energies related to orbitals? *[The principal quantum number gives a general indication of an electron's energy and sets restrictions on its orbital type. Higher energy orbitals are those with higher probability of being far from the nucleus.]*

6. How do cations form? *[If an electron absorbs energy in excess of the amount corresponding to the highest energy orbital in its electron cloud, it will be ionized, that is, leave the atom. The resulting positive ion left behind is a cation.]*

7. What determines the ionization energy of an element? *[The nuclear charge (the binding force's source) and the energy possessed by the electron (its energy level) combine (actually compete) to determine additional (ionization) energy required.]*

8. What determines the commonly formed cation of an atom? *[When a metal and nonmetal combine, the overall process is exothermic, largely due to the highly negative lattice energy involved; thus reaction takes place. One of the processes that requires energy in forming the compound is the ionization of the metal ion. It is the relative magnitudes of the ionization energies of a particular metallic species that determines the charge on the cation. For example, sodium has a first ionization energy (I_1) of 495.8 kJ / mol; its second ionization energy (I_2, the energy needed to remove a second electron from sodium) is 4562.4 kJ / mol, almost 10 times I_1! No reaction involving the formation of Na^{2+} could ever be expected to be exothermic; hence, Na^{2+} does not exist. A second example is the formation of magnesium ions. I_1, I_2, and I_3 for magnesium are 737.7, 1450.7, and 7732.6 kJ / mol, respectively. The large energy requirement for the formation of Mg^{3+} precludes its existence. Mg^{+} requires the least amount of energy for formation, but is unstable because of its half-filled 3s orbital. Hence, the most stable magnesium ion is Mg^{2+}, requiring a total energy expenditure of 2188.4 kJ / mol, which is certainly within the realm of a possible overall exothermic process for the formation of a magnesium salt.]*

9. What is the process of electron excitation? *[An electron is said to be excited if it absorbs energy and moves to a higher energy orbital.]*

10. Why do atoms emit light when excited? *[Being in an excited state means having excess energy. Like anything else, atoms tend to lose their excess energy, so the electron drops to a lower energy orbital. The "lost" energy is given off as light ($\Delta E = h\nu$).]*

11. How do the physical sizes of atoms and nuclei compare? *[Atoms tend to have diameters of around 0.1 nm (1-10 Å). The nucleus is about 10^{-15} m (1 femtometer) in diameter but contains 99.9% of the mass.]*

12. What are the three dimensional shapes of s and p orbitals? *[The 90% probability boundary for an s-orbital is a sphere. For p-orbitals the boundary is "dumbbell" shaped. Three p-orbitals, one dumbbell per axis, comprise a set.]*

13. What is meant by "wave-particle duality"? *[In some instances, the behavior of electromagnetic radiation and electrons can best be understood in terms of wave theory, and in other instances, their behavior can best be understood if we treat them as discrete particles. Electromagnetic radiation is usually described in terms of wavelength or frequency; the quantity most closely associated with particles is mass or momentum, the mass-velocity product. Louis de Broglie in 1924 offered the startling proposition that light may sometimes display particle-like characteristics, and that small particles may sometimes display wavelike properties. He summarized his proposition in a now famous equation $\lambda = h / mv$, where λ is the wavelength, h is the Planck constant, m is the mass, and v is the velocity of the object in question. Wave-particle duality is only important when the wavelength is of atomic dimensions,* i.e., *of the order of 10^{1} - 10^{3} picometers.]*

14. What information is provided by quantum numbers? *[All we can hope to know about an atomic electron: its energy, angular momentum, magnetic moment and spin.]*

15. How do quantized and continuous processes differ? *[Anything quantized can be described in terms of separate discrete units. Those that are continuous are without beginning or end and may be had in any quantity. Ice cubes are quantized but water flows continuously.]*

16. What is the electron configuration of (specify an element)?

Examples:	*H*	$1s^1$
	Li	$1s^2 2s^1$
	Ga	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$

Counterintuitive Examples

1. Adjustable light dimmer switches (unless specifically designed) do not work on fluorescent lamps. You can demonstrate this fact by first connecting a standard, preferably clear, incandescent light bulb to a dimmer switch or laboratory voltage controller such as a Powerstat™. By slowly increasing the voltage, you can brighten the incandescent bulb in a continuous manner. Repeat using a fluorescent lamp. Since the mercury gas enclosed in the fluorescent lamp's envelope requires a minimum ionization energy, a quantized process, the lamp will not go on until this minimum energy is applied.

2. A weight attached to a string and rotated around one's head continues in motion only as long as energy is applied. An electron, perceived to be in a classic Bohr orbit, does not slow down. This behavioral discrepancy was one of the counterintuitive problems that Bohr had to solve.

3. Energy changes do not always seem to be quantized. If you heat a piece of metal, its temperature rises on a continuum; the metal does not seem to acquire energy in quantum leaps. On the other hand, if you step halfway between two ladder rungs, you find that you remain on the first step; you must acquire potential energy in quanta dictated by the step-size of the ladder. The manner of acquiring energy in these two instances is not different. The energy step-sizes for a metal are so infinitesimally small that they cannot be detected by a relatively crude probe such as a thermometer.

4. As electrons are added to atoms in the same period of the Periodic Table, the atomic size decreases. Adding an electron also entails adding a proton to the nucleus, thus increasing effective nuclear charge. The nuclear charge density increase is proportionally greater than the electron energy increase, leading to more compact (smaller) atoms. A Periodic Table showing atomic sizes (such as the one included in the *Periodicity* module) will help illustrate this point.

5. Light radiated by an excited atom consists of several colors, but the human eye perceives only one composite color. This composite color can be resolved into its components by viewing it through a diffraction grating. A spectral tube for neon or helium, or even an orange night light, illustrates this fact very nicely.

6. Because electron probability functions are asymptotic, they do not become zero even at a great distance from the atom's nucleus. Consequently, there is a very small chance that some electron associated with an atom in your body is at this moment very far away. Perhaps a part of you has been to the moon and you just don't know it!

Analogies and Metaphors

1. **Size relationships of subatomic particles**
    - a. If the nucleus were the size of a period (.) the atomic diameter would be about 5 m. (The volume of the nucleus is about 1/100,000 the volume of the entire atom.)
    - b. If the nucleus of an atom were the size of a baseball, the atomic diameter would be about 4 km. The electrons would each be smaller than a period (.) and would move about at random in the spherical region between the nucleus and the edge of the atomic diameter.
2. **Sizes of atoms**
    - a. The approximate diameter of an atom is 10^{-8} cm. A speck 0.1 mm in diameter, about half the size of the period at the end of this sentence, requires one million atoms.
    - b. There are approximately 100 billion billion ($100 \times 10^9 \times 10^9$) atoms in a tiny drop of water.
    - c. It would require a million atoms, edge to edge, to match the thickness of a page of paper.
    - d. Comparing atomic diameter to one centimeter. If atoms were placed side by side, 100 million atoms would stretch 1 cm.

      100,000,000 atoms
      •••••
      | 1 cm |
3. **Atomic mass**

    Relative masses:

Proton	1 u
Electron (1,835 electrons = Mass of 1 proton)	0.00055 u
120 lb. person	32,770,000,000,000,000,000,000,000,000 u

4. **Density of nucleus**

    Because an atom is mostly space, few people realize the denseness of the nucleus. The nucleus has a density of 1×10^{14} g/cm^3. An object with the volume of a penny made of this material would have a mass in excess of 30 million tons.

Pictures in the Mind

1. A difficult concept for students is the relation between nuclear size (10^{-13} m) and atomic size (10^{-8} m). A good mental picture of this comes from the analogy of a grain of sand on a football field. Divide students into small groups. Give each group a round object to represent an atom's nucleus (orange, tennis ball, marble, *etc.*). Tell them to calculate how large the corresponding atom would need to be to correspond to the scale of the nucleus.
2. Use balloons to represent an *s*- or *p*-orbital.
3. A good analogy or picture in the mind for quantization is egg cartons or cans of soda from a machine (*i.e.*, if one can costs 50¢, even though you have 75¢ you can only get one can). Another example is a bookcase. Rough-sketch a bookcase on the board; show how a book can be placed on the second shelf, or on the third shelf, but it cannot be placed halfway between. Placement of books is quantized. Or, alternatively, borrow a ladder from the school custodian. Have a student climb up to the second step; then to the third step. Then ask the student to stand *midway* between Steps 2 and 3. A ladder represents quantized levels, while a ramp represents a continuum.

4. A Hund's rule analogy might be an empty bus that has seats, each holding two persons. As people file onto the bus, they usually occupy a seat singly until each seat has one person in it. Then, as additional people enter the bus, they begin to double up in the seats.
5. Set up three pairs of chairs in three rows. Ask three students (same sex works best) to sit anywhere in three of the chairs. With a high regularity, the distribution will be one person per pair of chairs. They will likely sit facing the same direction. This is analogous to electrons distributing one per orbital (with parallel spins) when the orbital energies are identical.
6. As an analogy for the wave/particle duality of electrons, display separate samples of whole peanuts (representing particles), creamy peanut butter (representing continuous waves), and chunky peanut butter (representing the dual nature).
7. Quantum numbers can be thought of as similar to zip codes. In a zip code, the first number indicates a geographical region of the U.S.A., while the subsequent numbers refer to more and more specific locations, like state and town. Similarly, quantum numbers begin with the general location of the electron, n, and progress to the specific orientation, m_s.

	General → Specific
Zip Code	92093
Quantum Numbers	n l m_l m_s 3 1 −1 +1/2

8. To get an idea of how dense the nucleus of an atom is, try this thought experiment: imagine the earth compacted to the size of a thimble. The resulting density roughly represents the density of the nucleus.

Problem Solving

1. To encourage understanding of pertinent mathematical relationships, ask students to rewrite the equations $c = \lambda\nu$ and $E = h\nu$ in all possible forms. Then have them express in words (not symbols) what each form of the equation "says." *[$\lambda = c/\nu$ "The wavelength is equal to the speed of light divided by the frequency." $\nu = c/\lambda$ "The frequency is equal to the speed of light divided by wavelength. In all cases λ and ν are inversely proportional—the larger the λ the smaller the frequency . (Same manipulations for $E = h\nu$ equation.)]*
2. Encourage students to analyze the value of E for appropriate units. When E has a small value, it is expressing energy *per photon*, rather than energy *per mole* of photons. This distinction is frequently overlooked in textbooks.

TIPS FOR THE TEACHER

Language of Chemistry

atomic mass (u) mass of an atom, expressed in atomic mass units.

atomic number *(Z)* number of protons in the nucleus of an atom.

atomic weight average atomic mass of all naturally occurring isotopes of an element, expressed in atomic mass units (u or amu).

Atomic spectrum includes the electromagnetic radiation produced by *emission*, due to electrons falling from an excited state to a lower energy state, or light that is *absorbed*, due to the raising of electrons to a higher energy state.

aufbau principle scheme used to reproduce the electronic configuration of multi-electron atoms. Shells and subshells are filled successively from low energy to higher energy in a specific order, according to stability.

electromagnetic spectrum range of wavelengths and frequencies of radiation including visible, ultraviolet, and infrared light, as well as X-ray and microwave radiation.

energy (E) ability to do work.

energy level specific values of energy in an atom.

frequency (ν) number of wavelengths per second that pass a fixed point.

ground state lowest energy state possible for an atom. Energy states other than the ground state are called *excited states*.

Heisenberg's uncertainty principle qualitatively, this principle states that one can never know both the exact position and momentum of an electron at the same time.

Hund's rule in the ground state electron configuration, all orbitals of equal energy are first occupied by single electrons with parallel spins before any orbital is doubly occupied.

isotope atoms with nuclei containing the same number of protons but different numbers of neutrons.

light part of the electromagnetic spectrum that is visible to the human eye.

mass number integer referring to the sum of all protons and neutrons in an atom's nucleus.

nucleus central portion of the atom containing protons and neutrons. The nucleus contains most of the mass of an atom, but less than 0.01% of its volume.

orbital region of space where there is a high probability of finding an electron. Mathematically, this is described by a wave function.

Pauli exclusion principle states that no two electrons in an atom may have the same four quantum numbers. It therefore follows that any orbital can contain only two electrons of opposing spins.

photon particle or quantized packet of electromagnetic radiation with an energy, wavelength, and frequency.

Planck's constant (h) proportionality constant relating the energy of a photon to its frequency. $h = 6.63 \times 10^{-34}$ J s.

quantized having only discrete values possible.

quantum numbers four quantum numbers are assigned to each electron in an atom. The *principal quantum number, n,* identifies the shell where the electron can be found; the energy of the electron mainly depends on the principle quantum number. The *angular momentum quantum number, l,* distinguishes the subshells within each shell. Each value of l corresponds to a subshell of different shape. The *magnetic quantum number, m_l,* indicates the directional orientation of orbitals within a subshell. Lastly, the *spin quantum number, m_s,* refers to the two possible orientations of the electron's spin axis. Any two electrons in the same orbital must have opposing spins.

ROY G BIV acronym for the order of colors in the visible spectrum: Red, Orange, Yellow, Green, Blue, Indigo, Violet.

shell the principal energy level occupied by an electron; indicated by the *principal quantum number*.

subshell indicated by angular momentum quantum number, *l*, this corresponds to energy sublevels within the principal quantum shell. *s*- and *p*-subshells are examples.

wavelength (λ) the distance between successive peaks or troughs in a wave.

Note on Electron Configuration

Writing electron configurations for the first 20 elements is straightforward. Orbitals written in the order in which they fill correspond to the numerical order in which the orbitals occur. For example, calcium (Z = 20) has a configuration of $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$, or in shorthand notation, $[Ar]4s^2$. With the next element, scandium (Z = 21), we must make a decision since either the numerical order or the order of orbital filling will be broken—they do not correspond. The configurations for scandium can be written as: (a) $[Ar]3d^1 4s^2$ or (b) $[Ar]4s^2 3d^1$. Both representations are commonly used and both are correct. In (a), all the subshells of a principal shell are grouped together, and the subshells of highest principal quantum number are placed last. In (b), the subshells are written in apparent order of electron filling. In *Sourcebook*, we have chosen to write electron configurations according to the expression in (a). This method has two advantages: (1) it keeps the subshells of each principal quantum number together, and (2) it is easy to see which electrons will be lost first upon chemical reaction (those of the highest principal quantum number, which are placed last).

Units

- Centimeter (cm) - 1/100 of a meter; 100 cm = 1 m
- Nanometer (nm) - One trillionth of a meter; 10^9 nm = 1 m
- Hertz (Hz) - Cycles or events per second
- Reciprocal seconds (s^{-1})
- Kilojoules (kJ) - 1000 joules. The joule is the SI energy unit. 4.18 kJ = 1 kcal
- Mole (mol) - Amount of substance that contains 6.02×10^{23} elementary entities, such as atoms, molecules, electrons, photons, *etc.*

Pattern Recognition

Teaching Electron Configuration (The Aufbau Process)

n, n^2, $2n^2$ Relationships

n	Subshells	n^2	$2n^2$
1	1s	1	2
2	2s2p	4	8
3	3s3p3d	9	18
4	4s4p4d4f	16	32
5	5s5p5d5f5g	25	50

n = Principal quantum number (energy level)
n = Number of kinds of subshells at energy level n
n^2 = Number of orbitals (total) at energy level n
$2n^2$ = Maximum number of electrons at energy level n

To "see" the pattern of electron configuration in atoms and "read" it by looking at the Periodic Table of Elements, consider the following method. First, use an atomic energy-level diagram to introduce the relationship between excitation and emission. Then, look at or draw the energy-level diagram for many-electron atoms (see Figure 10).

Introduce students to the n, n^2, $2n^2$ relationships for electrons in an energy level. Place this information onto the energy-level diagram.

ENERGY LEVEL DIAGRAM

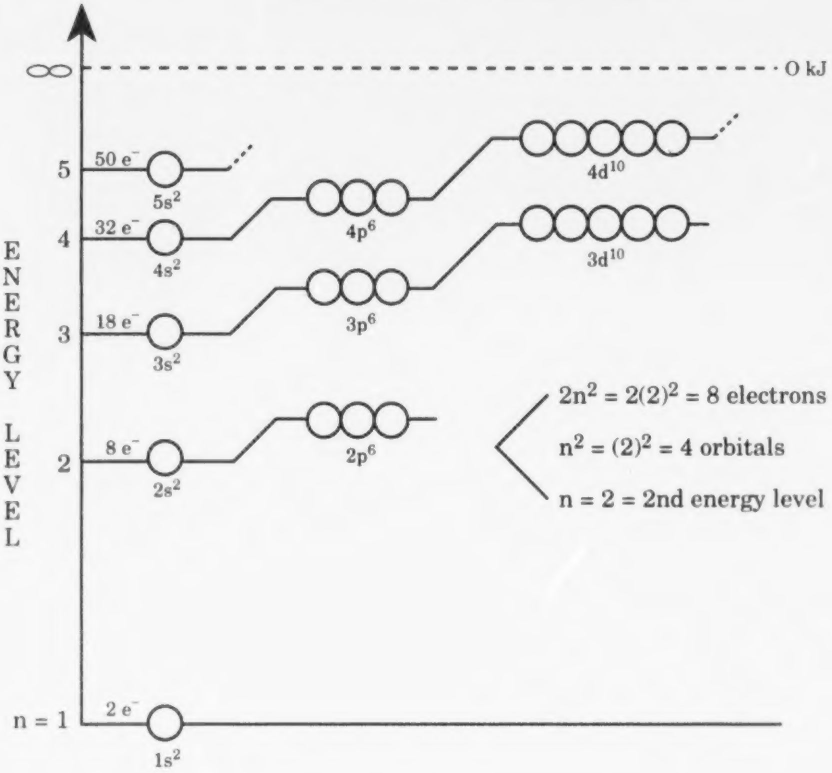

∞
0 kJ
ENERGY LEVEL
5
4
3
2
n = 1
50 e⁻
32 e⁻
18 e⁻
8 e⁻
2 e⁻
$5s^2$
$4s^2$
$3s^2$
$2s^2$
$1s^2$
$4p^6$
$3p^6$
$2p^6$
$4d^{10}$
$3d^{10}$
$2n^2 = 2(2)^2 = 8$ electrons
$n^2 = (2)^2 = 4$ orbitals
n = 2 = 2nd energy level

Figure 10. Atomic energy level diagram.

The maximum number of electrons in each energy level can now be determined: n = 1, ($2\,e^-$); n = 2, ($8\,e^-$); n = 3, ($18\,e^-$); n = 4, ($32\,e^-$); *etc.* The maximum number of electrons per subshell can also be found: s = $2\,e^-$, p = $6\,e^-$, d = $10\,e^-$, f = $14\,e^-$, *etc.*

Next, have students draw a Periodic Table showing the first 36 elements (H → Kr). Show the atomic number and symbol for each element. Now, following the order of the energy-level diagram, write the complete electron configuration for each element below its symbol. After this exercise, have students look for both a horizontal and vertical pattern to the outer (valence) electrons and the position of elements in the Periodic Table.

s = $2\,e^-$
p = $6\,e^-$
d = $10\,e^-$
f = $14\,e^-$

Maximum of $2\,e^-$ per orbital

What is noted in different rows and columns that shows a pattern to the order of electron build-up from lowest to highest energy level? (The order is called the *Aufbau process.*) If students don't copy from a book, they will make two errors in the electron configuration of the first 36 elements: Cr and Cu do not follow this pattern. These two elements are exceptions; their electron configurations must be memorized, if needed.

Periodic Table Pattern to Electron Configuration
The Aufbau Process

Elements-Atomic Number 1-36			
$1s^2$			
$2s^2$			$2p^6$
$3s^2$			$3p^6$
$4s^2$	$\underline{3}d^{10}$		$4p^6$
Elements-Atomic Number 37-118			
$5s^2$	$\underline{4}d^{10}$		$5p^6$
$6s^2$	$\underline{5}d^{10}$	$\underline{\underline{4}}f^{14}$	$6p^6$
$7s^2$	$\underline{6}d^{10}$	$\underline{\underline{5}}f^{14}$	$7p^6$

Associate rows in the Periodic Table with electrons filling the same energy level as the row: first row Æ first energy level; 1s; second row Æ second energy level 2s2p; *etc.*

$s^1 \rightarrow s^2$ $\qquad$ $s^2p^1 \longrightarrow s^2p^6$

$s^2d^1 \rightarrow s^2d^{10}$

La

Lanthanide Series Ce to Lu $\mathbf{6s^2\ \underline{5}d^1\ \underline{\underline{4}}f^{1 \rightarrow 14}}$

Ac

Actinide Series Th to Lr $\mathbf{7s^2\ \underline{6}d^1\ \underline{\underline{5}}f^{1 \rightarrow 14}}$

Notice that d-filling subshells fill one energy level below the outer s-subshell. The f-filling subshells fill two energy levels below the outer s-subshell.

Relate specific groups (families) to the kinds of subshells that electrons are filling.

The two inner-transition series for La and Ac have many exceptions. Compare this pattern to the presentation in your chemistry text.

The *Aufbau Process* and Electron Configuration

$_{Z}X\ \ 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 \underline{3}d^{10} 4p^6 5s^2 \underline{4}d^{10} 5p^6 6s^2 \underline{5}d^{10} \underline{\underline{4}}f^{14} 6p^6 7s^2 \underline{6}d^{10} \underline{\underline{5}}f^{14} 7p^6$, *etc.*

Common Student Misconceptions

1. **"Electrons orbit the nucleus like the planets around the sun."**

   This incorrect model of the atom can still occasionally be found in popular science literature or in old reference books. If it arises, students should be told that it is a poor model, but early in this century, in the early development of atomic theory, it was an important starting point in the acceptance of quantized energy for electrons in atoms. This model is also called the Bohr model; one might also refer to Bohr atoms.

2. **"Each orbital exists alone in space."**

   Students fail to recognize that in an n = 2 shell, the 2*s*, and all three 2*p*-orbitals are superimposed in space and have about the same size. They also don't recognize that the different energy levels (n = 2 and n = 1 for instance) exist in the same space.

3. **"Radiometers (hollow glass globes with a free-moving vane inside) prove that photons of light have mass."**

   Indeed, when a light shines on the vane, it rotates nicely. Actually, there is a small amount of air in the bulb. Air molecules on the black side of the vane's projections warm up more quickly than those on the shiny side. The warmed gas molecules move faster, pushing against the vane, so it moves.

4. **"An orbital is the same as an orbit."**

   Orbits (associated with the planetary Bohr atom model) specify fixed trajectories for electron movement. An orbital, representing a three-dimensional region within which one or two electrons can be found, provides no information regarding specific electron movement.

5. **"There is an edge, or boundary, to an atom."**

   Atoms and orbitals have no "hard" edges. Instead, the contour of the orbital, as we show it in pictures and models, corresponds to the space of highest probability for electron location.

6. **"Air exists between particles (protons, neutrons, and electrons) in atoms."**

   There is no matter, air or otherwise, between atomic particles, since they *are* the matter.

7. **"The electromagnetic spectrum extends from red to violet light."**

   Light is the portion of the electromagnetic spectrum visible to the human eye, but the entire spectrum includes X-rays, radio waves, infrared radiation, and ultraviolet radiation, as well.

HISTORY: ON THE HUMAN SIDE

1. **Dmitri Mendeleev (1834 - 1907)**

   During the mid-1860s, Mendeleev was stimulated by the abolition of serfdom in Russia and began concerning himself with the practical problems facing his nation. In 1865 he held a series of lectures at the Free Economic Society on experimental agriculture, the cooperative production of cheese, and techniques for agricultural fertilization. Beginning in 1867, Mendeleev conducted research on the relation and organization of elements. He established correct atomic weights and arranged the elements in a periodic chart that left blanks for unknown elements (such as helium, neon, argon, gallium, and germanium, which were discovered later). The discovery of gallium in 1875 and germanium in 1886, both of which showed properties predicted by the model, resulted in broad acceptance of Mendeleev's periodic table by the scientific community.

2. **Max Planck (1858 - 1947)**

   While a student at the University of Munich, Planck carried out his only laboratory experiments, such as on osmosis of gases. All of his subsequent work was in the area of theoretical physics/chemistry. During the 1890s he began doing research in electrodynamics, and he soon became interested in heat and "black body" radiation. In 1899 he completed a derivation of the special distribution law for heat radiation and examined the difference between absorbed and emitted energy of bodies. The equations he published formed the basis of the radiation formulas that describe the quanta (packets of energy) emitted by heated bodies. Planck was awarded the Nobel Prize in physics in 1918 and continues to be known as the founder of quantum theory. In the later years of his life, Planck began writing on the history and philosophy of science and religion. He examined the divisions between theoretical and applied physics and the role of religion and science in waging a timeless battle against skepticism and superstition.

3. **Albert Einstein (1879 - 1955)**

   Einstein's dissertation entitled "A New Determination of Molecular Dimensions" revolutionized chemistry and physics by hypothesizing that light should be considered a collection of independent particles of energy, called light quanta. The wave theory of light no longer accounted for the experimental results of blackbody radiation. Einstein's research showed that the energy (E) of particles was proportional to the frequency (ν) of the radiation: $E=h\nu$. This formula accounted for the properties of fluorescence as well as the photoelectric effect. In 1907, as a result of further research in this field, he formulated the famous equation $E = mc^2$, which formed the foundation of the relativity theory and much of 20th century physics. Once this ground-breaking work was recognized, Einstein was offered a variety of prestigious positions before moving to Berlin in 1914 as a member of the Prussian Academy of Sciences, where he had colleagues such as Max Planck and Erwin Schrödinger.

   During WWI, Einstein's research culminated in the general theory of relativity. Once the term relativity had become a household word, Einstein used his international position to advocate pacifism. When Hitler came to power in 1933, Einstein resigned his position in the Berlin Academy of Sciences and moved to Princeton, New Jersey to work at the Institute for Advanced Study. He remained there for the remainder of his life, becoming a U.S. citizen in 1940. Upon recognizing the devastating results of the atomic bomb Einstein, devoted much of his energy to pacifism and the establishment of the United Nations.

4. **Werner Heisenberg (1901 - 1976)**

   In 1924 Heisenberg went to the Institute for Theoretical Physics in Copenhagen to study under Niels Bohr who had developed the planetary model of the atom. According to this model, electrons move in well-defined orbits around the nucleus. However, by 1925 this theory was no longer reconcilable with experimental evidence, and the quantum mechanical model was beginning to evolve. Influenced by Einstein's work on relativity, Heisenberg began using variables to represent various observable quantum conditions. His experiments on atomic and molecular spectra, ferromagnetic phenomena, and electromagnetism led him in 1929 to state his famous "uncertainty principle." This principle describes a fundamental limitation of modern physics—one cannot determine with accuracy *both* the momentum and position of a particle at the same time. After WWII he organized and became director of the renowned Max Planck Institute for Physics and Astrophysics in Göttingen and moved to Munich in 1958 when the Institute relocated there.

5. **Niels Bohr (1885 - 1962)**

   Niels Bohr and Erwin Schrödinger played important roles in developing theories on atomic structure. Most textbooks document their achievements.

HUMOR: ON THE FUN SIDE

1. Mnemonic aid to recall the order of orbitals in order of increasing energy:

   _S_ome _P_eople _D_on't _F_orget.

2. Sign on bumper stickers:
   a. Heisenberg may have slept here.
   b. Wave if you've met Schrödinger.
   c. Be discrete—use quanta.

3. Student responses on exams:
   - a. Atoms can expand and travel.
   - b. An atom is a discrete part of a particle.
   - c. Atoms are composed of exceedingly minute particles called elements.

4. **IONS MINE**

We are all familiar with J. J. Thomson's reputation as a physicist but few recognize his talent for song writing. This song, with his lyrics, is to be sung to the tune of "Darling Clementine."

1. In the dusty lab'ratory,
'Mid the coils and wax and twine,
There the atoms in their glory,
Ionize and recombine.

Chorus: Oh my darlings! Oh my darlings!
Oh my darling ions mine!
You are lost and gone forever
When just once you recombine!

2. In a tube quite electrodeless,
They discharge around a line,
And the glow they leave behind them
Is quite corking for a time.

3. And with quite a small expansion,
1.8 or 1.9,
You can get a cloud delightful,
Which explains both snow and rain.

4. In the weird magnetic circuit
See how lovingly they twine,
As each ion describes a spiral
Round its own magnetic line.

5. Ultra-violet radiation
From the arc of glowing lime,
Soon discharges a conductor
If it's charged with minus sign.

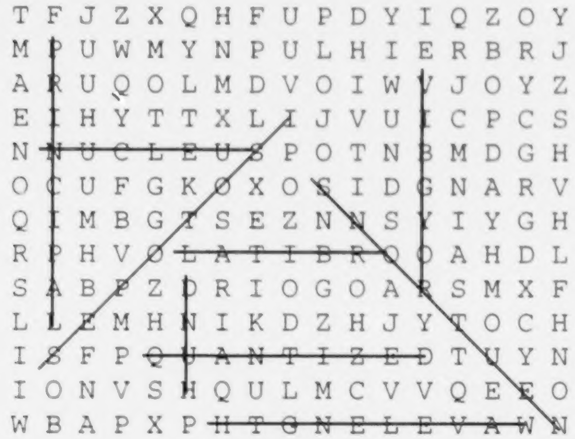
6. Alpha rays from radium bromide
Cause a zinc-blende screen to shine,
Set it glowing, clearly showing
Scintillations all the time.

7. Radium bromide emanation,
Rutherford did first divine,
Turns to helium, then Sir William
Got the spectrum—every line.

[Chem 13 News/March 1981/p. 3]

5. Tom Weller, Science Made Stupid, Houghton-Mifflin Co., Boston, 1985, pp. 23, 25,75. This humorous paperback pokes fun at atomic theory, dinosaurs, evolution, and the universe in general.

6. Word Search (see *Appendix* for master copy)



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T F J Z X Q H F U P D Y I Q Z O Y
M P U W M Y N P U L H I E R B R J
A R U Q O L M D V O I W V J O Y Z
E I H Y T T X L I J V U I C P C S
N N U C L E U S P O T N B M D G H
O C U F G K O X O S I D G N A R V
Q I M B G T S E Z N N S Y I Y G H
R P H V O L A T I B R O O A H D L
S A B P Z D R I O G O A R S M X F
L L E M H N I K D Z H J Y T O C H
I S F P Q U A N T I Z E D T U Y N
I O N V S H Q U L M C V V Q E E O
W B A P X P H T G N E L E V A W N
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Difference in the mass number and the number of protons is equal to the number of ____.
2. Region of space where there is a high probability of finding an electron.
3. Quantum number that identifies the shell where the electron can be found.
4. Acronym for the order of colors in the visible spectrum.
5. Where protons and neutrons are found.
6. Atoms with the same number of protons but different numbers of neutrons.
7. Having only discrete values possible.
8. Distance between successive peaks in a wave.
9. He stated that all orbitals of equal energy are first occupied by single electrons before any orbital is doubly occupied.
10. Number of protons in an atom's nucleus is symbolized by this letter.

Answers: 1. NEUTRONS 2. ORBITAL 3. PRINCIPAL 4. ROYGBIV 5. NUCLEUS 6. ISOTOPES 7. QUANTIZED 8. WAVELENGTH 9. HUND 10. Z

7. See cartoons at end of module.

MEDIA

1. CHEM Study Film, *Hydrogen As Viewed by Quantum Mechanics, Advanced Version*. Available from Ward's Natural Scientific, Inc., P.O. Box 92912, Rochester, NY 14692-9978. (800) 962-2660
2. Spectrum wall chart (available from Fisher Scientific Co., Educational Materials Division, 4901 LeMoyne Street, Chicago, IL 60651), with line spectra of many elements. This can be used to identify the wavelengths of lines observed in discharge tubes.
3. *Electron Configuration* film from EME Corporation, Old Mill Plain Road, Danbury, CT 06811.

4. Software published by *JCE:* Software, a publication of *the Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. *Rutherford*, by Robert Rittenhouse, Vol. V B, No. 2, for IBM PS/2, PC-compatible computers.
    b. *Animated Demonstrations*. by Philip Pavlik. Vol. V B, No. 2, for IBM PS/2. PC-compatible computers.
    c. *Hydrogen Atom Orbitals*, by Michael Liebl, Vol. III A, No. 2, for the Apple II computer.
    d. *Atomic Spectroscopy*, by Philip Pavlik, Robert Rittenhouse, Martin Rose and William F. Coleman, Vol. V N, No. 1, for IBM PS/2, PC-compatible computers.
    e. *Alpha Scatter*, by Robert Rittenhouse. One of the programs in The One-Computer Classroom, Vol. I A, No. 1 for the Apple II computer.
5. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    a. For the Apple II computer: AP 204
    b. For the Apple II computer running on ProDOS: AR 201, AR 202
    c. For IBM PCs and PC-compatibles: PC 2401
6. Videodiscs published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
    a. "The Rutherford Experiment," a chapter on *The World of Chemistry: Selected Demonstrations and Animations:* Disc I (double sided, 60 min.), Special Issue 3.
    b. "Electron Clouds" and "The Aufbau Principle," two chapters on *The World of Chemistry: Selected Demonstrations and Animations:* Disc II (double sided, 60 min.), Special Issue 4.
7. *Chemical Bonding and Atomic Structure*, 23 min. video available from Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).

INSTRUMENTATION

Diffraction gratings (available from Edmund Scientific, Central Scientific, or Sargent-Welch Scientific Company)

H_2, He, Ne, Hg gas discharge (spectrum) tubes

Power supply for tubes, or Tesla coil

Hand-held spectroscopes (available from Central Scientific)

Links/Connections

WITHIN CHEMISTRY

The outer-shell electron configuration of an atom determines the chemical behavior, physical behavior, and bonding of that element. Therefore this topic is the perfect lead-in for *Periodicity* or *Chemical Bonding*.

Before students can understand topics in *Nuclear Chemistry* they must know about the structure of the atom. So this topic unit is an important precursor.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Chemists make use of our knowledge at the atomic and molecular level to explain and predict the behavior of macroscopic materials. Some areas of physics concentrate on expanding knowledge at the sub-atomic level. Atomic structure is an area where chemistry and physics have a large overlap. Your school's physics teacher may have much of the laboratory and demonstration equipment needed for this topic.

Because an atom is a very small particle, it cannot be seen with the naked eye. Yet electron microscopes that can produce magnification factors in the millions have allowed scientists to photograph images of individual atoms. The first motion pictures of the movement of single uranium atoms was achieved in 1976 by physicists at the University of Chicago. What did an atom of uranium look like? The individual uranium atoms appeared as roundish spots with bright edges and dark centers.

Physicists continue to expand our notion of atomic structure by continually discovering new subatomic particles. Many are now known, *e.g.*, pions, muons, quarks, *etc*. How many more might yet be discovered when more powerful "atom smashers" are built? How might the super-collider presently being built in Texas help us to understand the atom better? (Idea contributed by Angie Matamoros.)

TO THE CONTEMPORARY WORLD

Television tubes are merely sophisticated cathode ray tubes. The electron stream from the rear of the tube is aimed at the front surface, which is coated with materials (called *phosphors*) that have readily excitable valence electrons. These electrons are temporarily energized. When they return to lower-energy states, the energy is released in the form of visible light with distinctive colors. Three major groups of phosphors (phosphorescent dots) in red, green, and blue arrays constitute one form of a color TV tube.

The details of atomic structure are particularly important in such medical applications as CAT, MRI, *etc*.

Extensions

Public Broadcasting System series *Creation of the Universe*, Stephen Hawking:

PBS Video
4401 Sunset Blvd.
Los Angeles, CA 90027

References

Module developed by Joe Baron, Marvin Blevins, and Barbara Sawrey, the California team.

Bouma, J. (1989). An application-oriented periodic table of the elements. *Journal of Chemical Education, 66,* 741-745.

This kind of periodic table provides an interest for finding the relevance of chemistry in everyday life.

Campbell, A. (1989). Let us make the table periodic. *Journal of Chemical Education, 66,* 739-740.

Describes a modified periodic table with symbols and data showing periodic properties.

Gribbin, J. (1984). *In search of Schrodinger's cat.* Bantam Books.

Kolb, D. (1977). But if atoms are so tiny... *Journal of Chemical Education, 54,* 543.

An explanation of how the masses of atoms are known.

Lemonick, M. D. (1990, April 16). The ultimate quest. *Time,* 50-56.

A contemporary essay on how physicists, armed with giant machines (electron-positron collider, linear collider, tevatron, and superconducting supercollider), spend billions of dollars in the race to discover building blocks of matter (such as leptons and quarks).

Peake, B. (1977). The discovery of the electron, proton, and neutron. *Journal of Chemical Education, 66,* 738.

Discusses the history of discovery and origins of the names of these three fundamental particles of an atom.

Ringnes, V. (1989). Origin of the names of chemical elements. *Journal of Chemical Education, 66,* 731-740.

Etymology and the reason scientist(s) coined a specified name for a newly discovered element are surveyed.

Roundy, W. (1989). What is an element? *Journal of Chemical Education, 66,* 729-730.

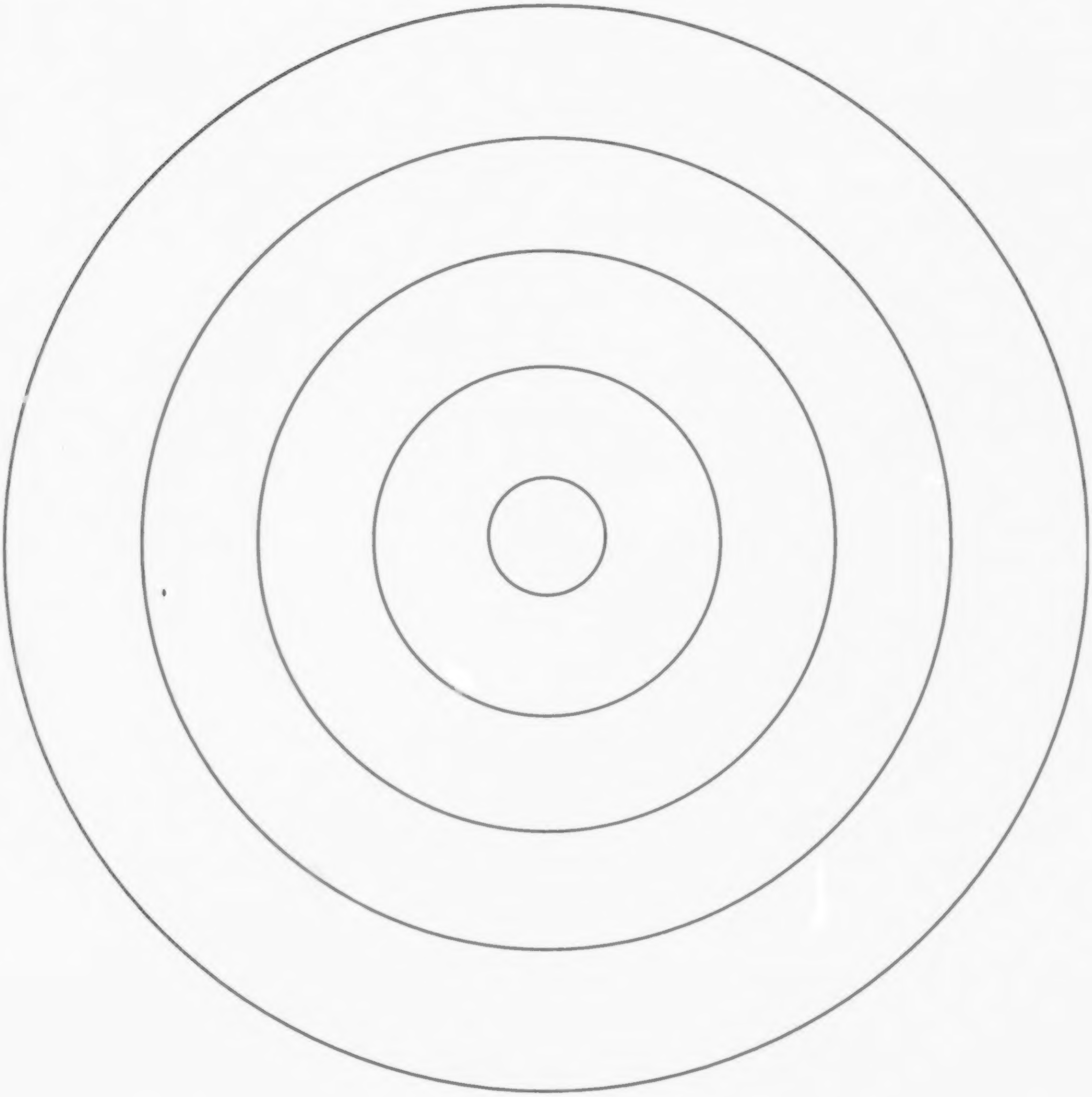
Each element is defined by the number of protons in the nuclei.

Zukov, G. (1979). *The dancing Wu Li masters.* Bantam Books.

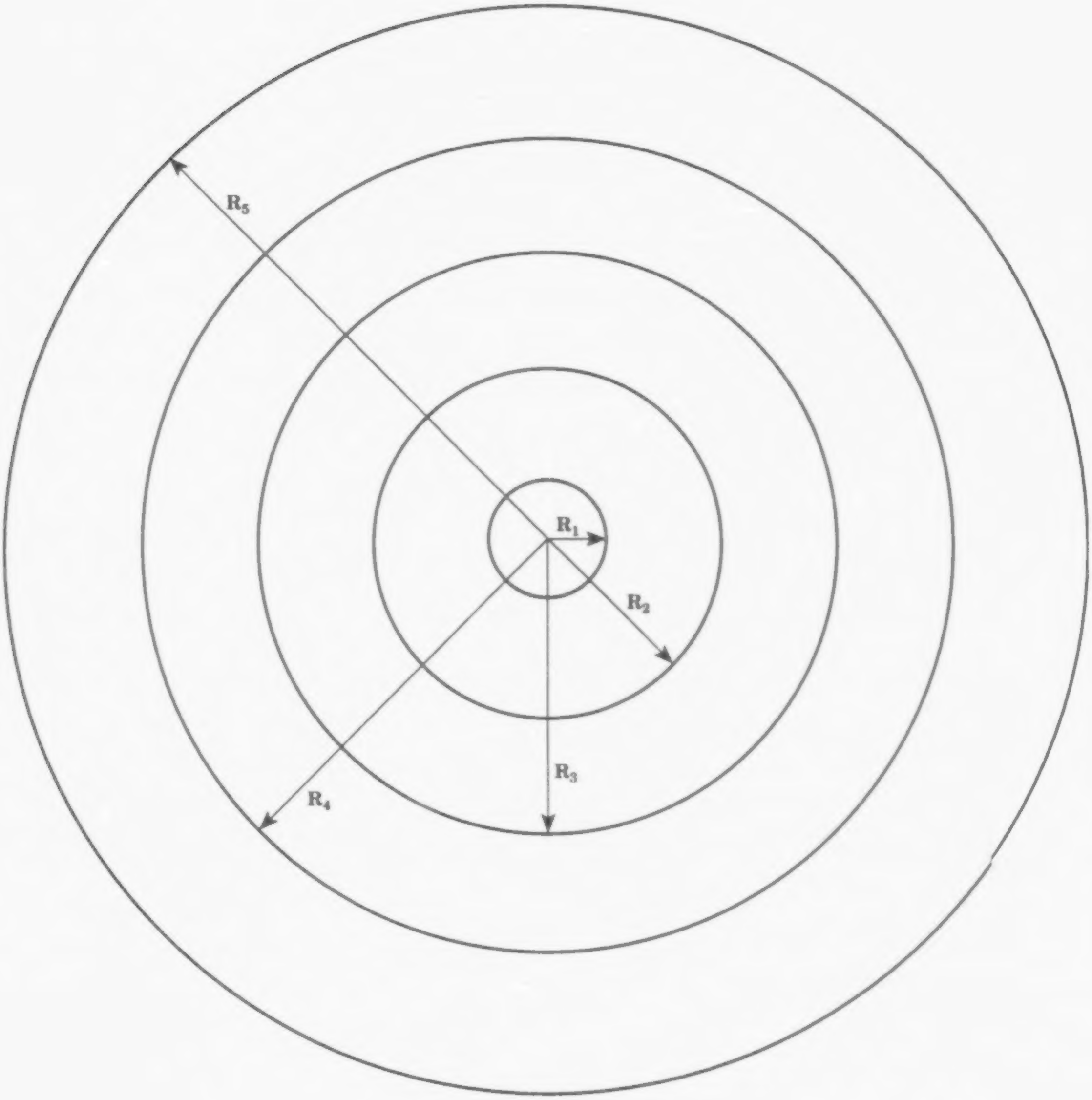
Appendix

- **Transparency Masters**
  1. Target for Activity 2
  2. Target for Activity 2 with Radii Shown
  3. Energy Level Diagram
  4. Laboratory-Aids #100 Obscertainer™ Kit
  5. Word Search
- **Humor**

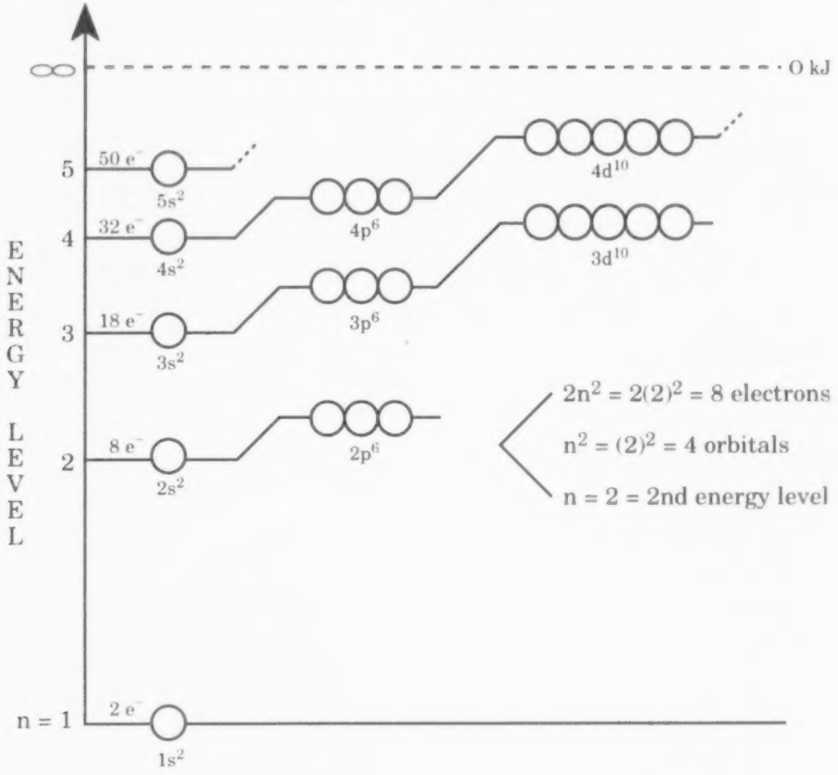
Target for Activity 2



Target for Activity 2 with Radii Shown


R5
R1
R2
R3
R4

ENERGY LEVEL DIAGRAM



∞
O kJ
ENERGY LEVEL
5
50 e⁻
5s²
4
32 e⁻
4s²
4p⁶
4d¹⁰
3
18 e⁻
3s²
3p⁶
3d¹⁰
2
8 e⁻
2s²
2p⁶
2n² = 2(2)² = 8 electrons
n² = (2)² = 4 orbitals
n = 2 = 2nd energy level
n = 1
2 e⁻
1s²

LABORATORY-AIDS® #100 OBSCERTAINER™ KIT

Student Worksheet and Guide

How many times have you looked at an object that attracted your attention and aroused your interest and asked the question, "What is it made of?" That same question has been asked thousands of times in many different languages. Because scientists continued to search until they found the answers to many of the problems they faced, we today have the modern world as our home.

We can look back at some of the answers suggested by the early Greek philosophers (as the scientists were called in those days) and smile. Today, even an elementary-school student knows, for example, that water is composed to two substances, hydrogen and oxygen. The Greeks, however, suggested that the world itself was made up of four elements of which water was one, and fire, earth, and air were the other three. Because of lack of careful observation and experimentation, people accepted the idea suggested by the Greeks. This acceptance is seen in our literary works where a storm, for example, is described as the raging of the elements. This idea was so well established that it was not until the seventeenth century that scientists established the fact that the Greeks were mistaken and science came closer to the answer it sought.

Most individuals consider themselves good observers. This laboratory activity involves many things, mainly, making careful observations and checking the accuracy of the hypothesis or best guess.

It is important for a science student to practice making useful observations even if they do not immediately lead to the complete solution of a problem. In this laboratory activity, you will begin by observing by indirect means. In order to solve the problem it takes concentration, alertness to detail, ingenuity, and patience.

Problem: What is the configuration (design) of the inside of a closed container (OBSCERTAINER)?

Theory: This laboratory activity is an exercise in indirect measurement or observation. The closed OBSCERTAINERS have partitions inside and a steel ball that can freely move. You will not be able to see or touch the inside of the OBSCERTAINER and yet you will have to determine the design of the inside by indirect means.

Procedure: Move the steel ball around by *carefully* shaking and tilting the OBSCERTAINER. By the sound and path of the steel ball, determine the shape and location of the partition(s). In the blanks below, indicate the identity of the OBSCERTAINER (its number) and your hypothesis. Retest your hypothesis and indicate any changes in the second drawing. This should reflect your final decision.

Some of the OBSCERTAINERS are more difficult than others, therefore time should be spent in making careful observations. Do not spend more than six minutes with each OBSCERTAINER. DO NOT OPEN THE OBSCERTAINER. DO AS YOUR INSTRUCTOR ADVISES. Study at least four different containers.

	HYPOTHESIS	RETEST	ACTUAL
OBSCERTAINER #	○	○	○
OBSCERTAINER #	○	○	○
OBSCERTAINER #	○	○	○
OBSCERTAINER #	○	○	○
OBSCERTAINER #	○	○	○

1. Write a summary of the actual OBSCERTAINER configuration and your hypothesis (conception) of it. List those things that you were able to determine and those that you were unable to determine.

2. Is there any reason why you were successful for certain characteristics and not for others?

Student's Name: **Date:**

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Word Search

```
T F J Z X Q H F U P D Y I Q Z O Y
M P U W M Y N P U L H I E R B R J
A R U Q O L M D V O I W V J O Y Z
E I H Y T T X L I J V U I C P C S
N N U C L E U S P O T N B M D G H
O C U F G K O X O S I D G N A R V
Q I M B G T S E Z N N S Y I Y G H
R P H V O L A T I B R O O A H D L
S A B P Z D R I O G O A R S M X F
L L E M H N I K D Z H J Y T O C H
I S F P Q U A N T I Z E D T U Y N
I O N V S H Q U L M C V V Q E E O
W B A P X P H T G N E L E V A W N
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Difference in the mass number and the number of protons is equal to the number of ____.
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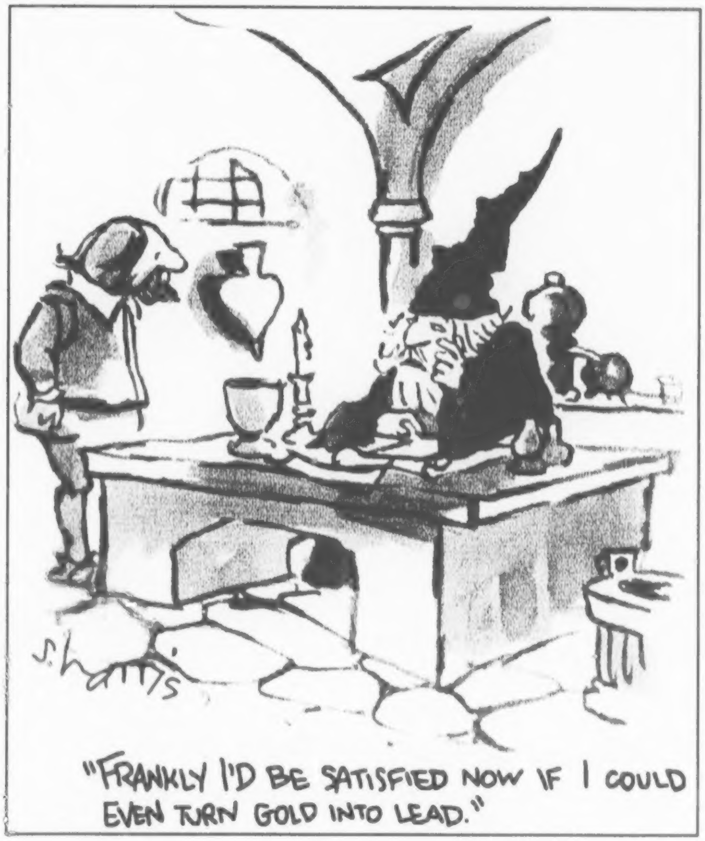
8. Distance between successive peaks in a wave.
9. He stated that all orbitals of equal energy are first occupied by single electrons before any orbital is doubly occupied.
10. Number of protons in an atom's nucleus is symbolized by this letter.


S. harris
"FRANKLY I'D BE SATISFIED NOW IF I COULD EVEN TURN GOLD INTO LEAD."

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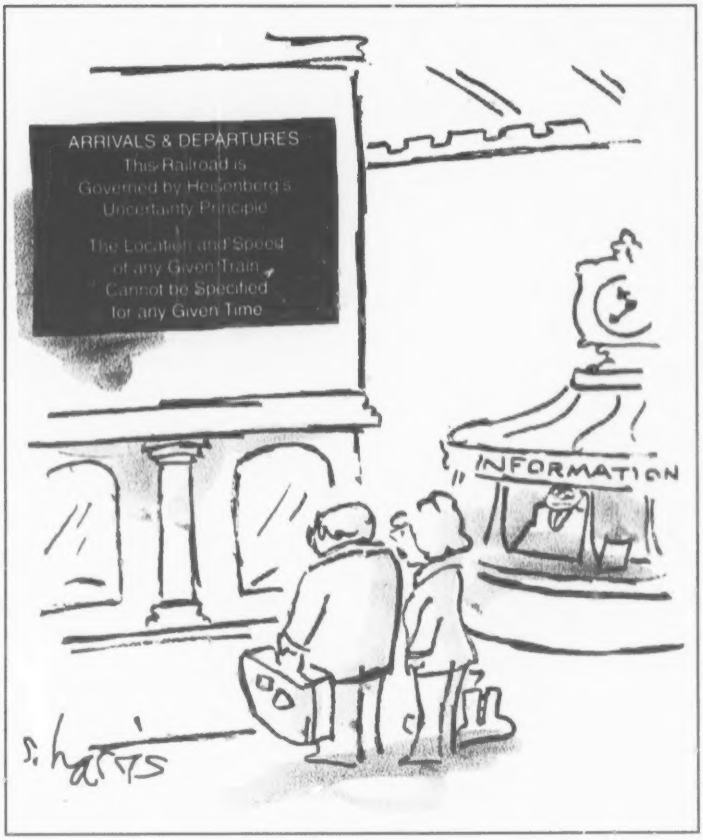
ARRIVALS & DEPARTURES
This Railroad is Governed by Heisenberg's Uncertainty Principle
The Location and Speed of any Given Train Cannot be Specified for any Given Time
INFORMATION
S. harris

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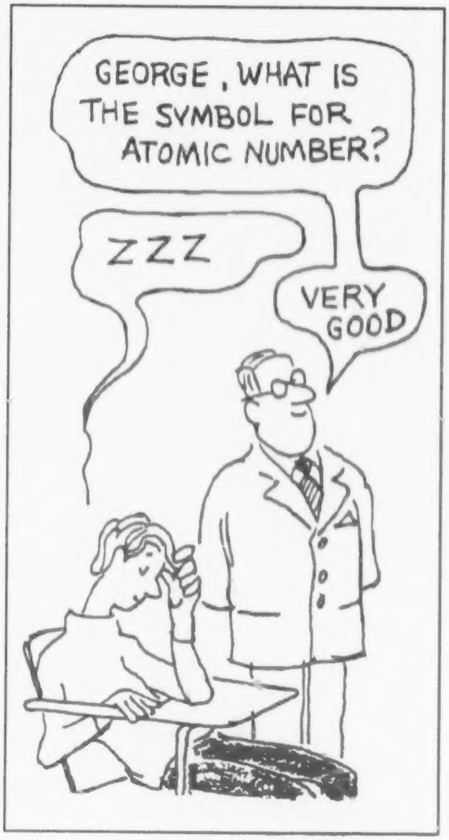
GEORGE, WHAT IS THE SYMBOL FOR ATOMIC NUMBER?
ZZZ
VERY GOOD

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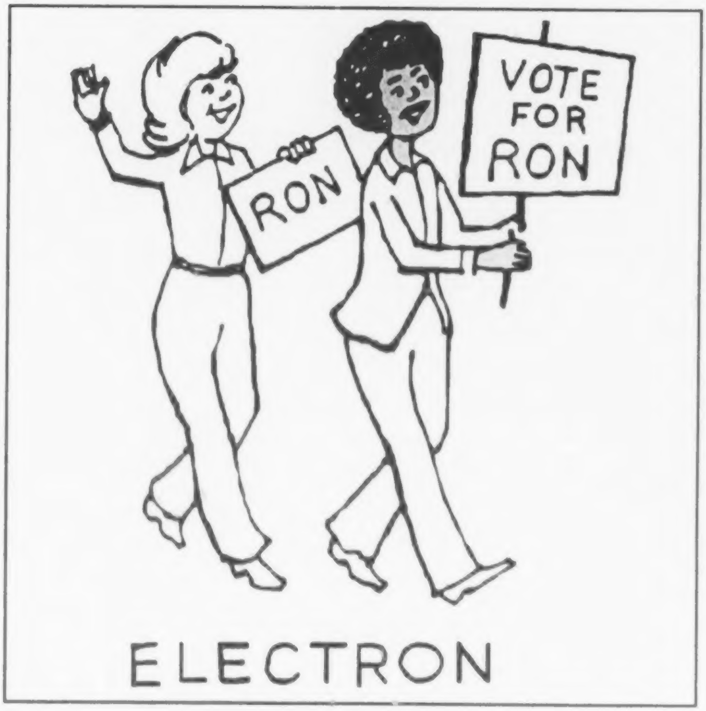
VOTE FOR RON
RON
ELECTRON

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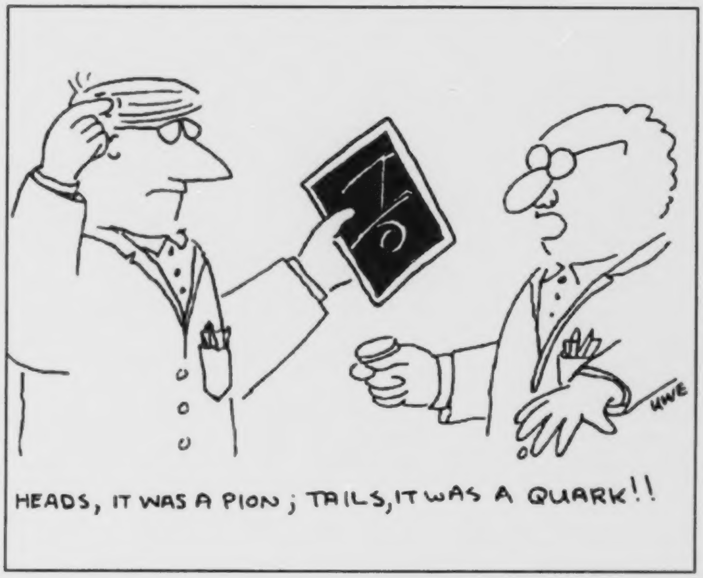
HEADS, IT WAS A PION; TAILS, IT WAS A QUARK!!
UWE

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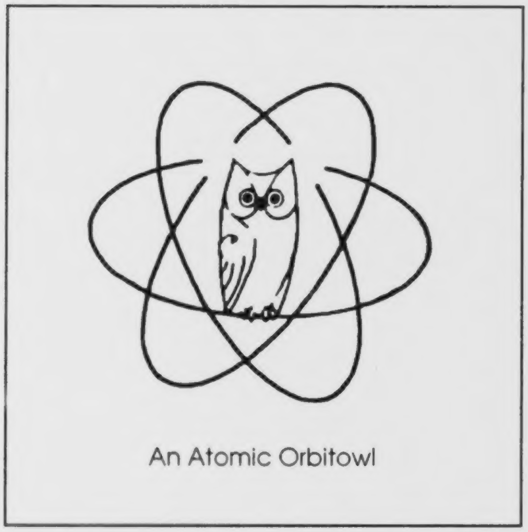
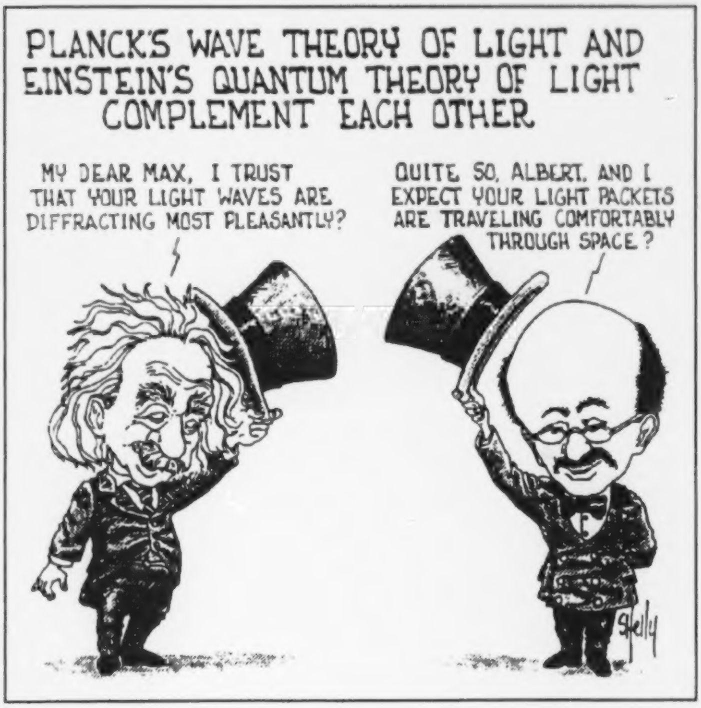
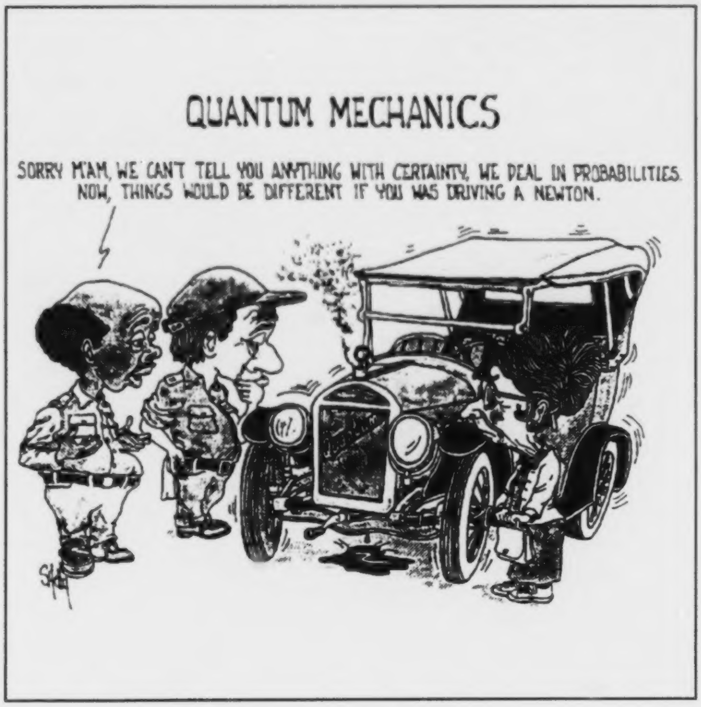

An Atomic Orbitowl

©CHEM 13 NEWS. (June 1970), p. 14. Reprinted with permission.


PLANCK'S WAVE THEORY OF LIGHT AND EINSTEIN'S QUANTUM THEORY OF LIGHT COMPLEMENT EACH OTHER
MY DEAR MAX, I TRUST THAT YOUR LIGHT WAVES ARE DIFFRACTING MOST PLEASANTLY?
QUITE SO, ALBERT, AND I EXPECT YOUR LIGHT PACKETS ARE TRAVELING COMFORTABLY THROUGH SPACE?
Shelly

Cartoon by Shelly Fischman. Reprinted with permission.


QUANTUM MECHANICS
SORRY M'AM, WE CAN'T TELL YOU ANYTHING WITH CERTAINTY, WE DEAL IN PROBABILITIES. NOW, THINGS WOULD BE DIFFERENT IF YOU WAS DRIVING A NEWTON.
1911

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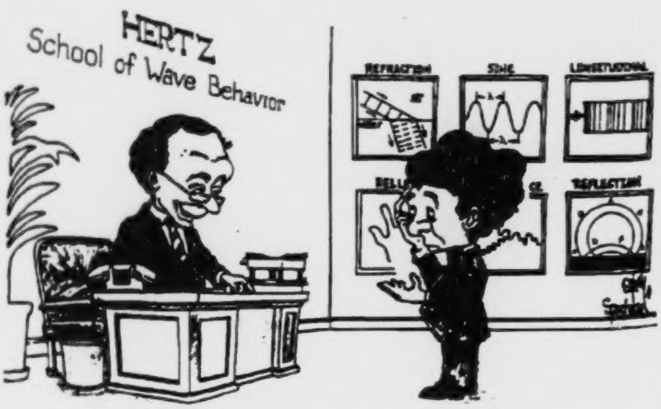
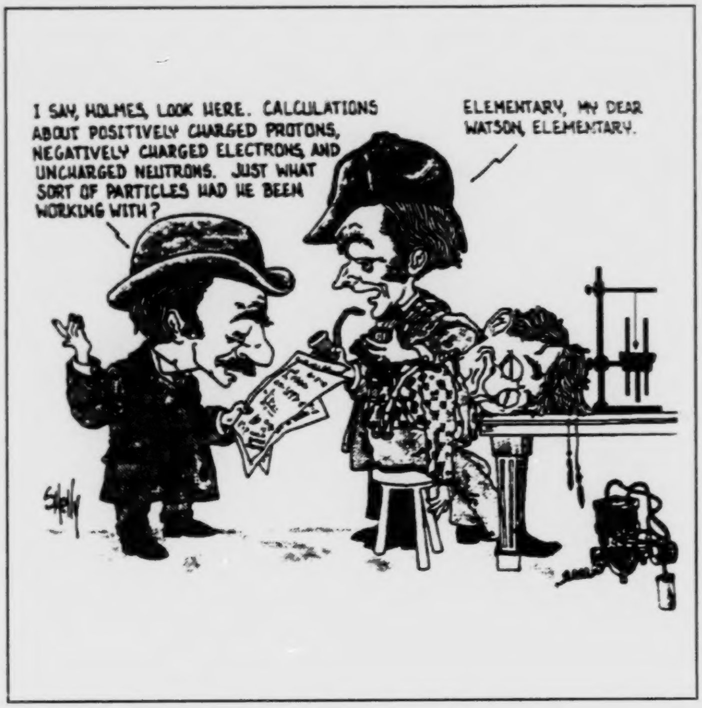

HERTZ
School of Wave Behavior
REFRACTION
SINE
LONGITUDINAL
REFLECTION

"DON'T WORRY, YOU'VE COME TO THE RIGHT PLACE"

Used with permission from McGraw-Hill Book Co.


I SAY, HOLMES, LOOK HERE. CALCULATIONS ABOUT POSITIVELY CHARGED PROTONS, NEGATIVELY CHARGED ELECTRONS AND UNCHARGED NEUTRONS. JUST WHAT SORT OF PARTICLES HAD HE BEEN WORKING WITH?
ELEMENTARY, MY DEAR WATSON, ELEMENTARY.
Shelly

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A SourceBook Module

Version 2.0 1994

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BIOGEOCHEMICAL CYCLES:
CARBON, OXYGEN, NITROGEN

ATMOSPHERE
HYDROSPHERE
LITHOSPHERE

ChemSource

Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Water and antifreeze circulate in automobile cooling systems to cool the engine. If the water level gets low, more water or antifreeze is added to allow the cooling cycle to operate. Unlike this example, however, the relative number of atoms of the chemical elements—upon which our lives depend—cannot be increased. The total quantity of the elements available to living organisms on earth is constant: the same atoms must be used repeatedly. This circulation of a fixed amount of the chemical elements between living organisms and the physical environment is a **biogeochemical** cycle.

There are several important features of biogeochemical cycles. First, every element has its own cycle, and the cycles are interrelated. This means, of course, that anything that interferes with the biogeochemical cycle of one element will alter the cycle of other elements. In this module, we will deal with three gaseous biogeochemical cycles: carbon, oxygen, and nitrogen. Second, there is usually more than one pathway available to an element or its compounds from one point to another in its cycle. If something should block one path, another path is usually available to take over. This is comparable to the collateral circulatory system of the heart after an infarction or a river channelling around an obstruction. Third, external factors can alter the natural cycles, producing adverse effects on the biota. Pollution of the environment is a primary factor that alters biogeochemical cycles today.

Since it is impractical to take an in-depth look at entire biogeochemical cycles, we will look at important selective parts of the cycles of carbon, oxygen, and nitrogen. As part of this investigation, we will conduct activities, perform demonstrations, and suggest additional assignments to look at the chemistry of specific steps in these cycles.

PLACE IN THE CURRICULUM

This module is enrichment material that could be taught with gases, acids and bases, organic chemistry or equilibrium. It may also be used after students have been exposed to a general first year chemistry course, since biogeochemical cycles afford an excellent opportunity to get an overall picture of real-world chemistry in action.

CENTRAL CONCEPTS

1. The total quantity of carbon, oxygen, and nitrogen available to all living organisms on earth is constant.
2. The elements circulate between living organisms and the nonliving, physical environment *via* biogeochemical cycles.
3. Earth's atmosphere is the product of an evolutionary process involving several on-going cycles.
4. Every element has its own cycle, although the cycles for all of the elements are interrelated.
5. Most biogeochemical cycles have several alternate pathways.
6. Coal formation is a form of geologic metamorphism.
7. Photosynthesis and respiration can be thought of as reverse reactions in terms of stoichiometry.
8. Human interventions can interrupt normal biogeochemical cycles. These "detours" include industrial processes, burning of fossil fuels and wood, deforestation, animal waste and artificial fertilizer run-off, use of pesticides, *etc.*

RELATED CONCEPTS

1. Nature of carbon and its reactions, producing inorganic carbon compounds.
2. Nature of oxygen and its reactions, producing compounds containing oxygen.
3. Nature of nitrogen and its reactions, producing nitrogenous compounds.
4. Stratospheric ozone and its absorption of specific solar electromagnetic radiation (wavelength and frequency).
5. Structure and function of carbohydrates, proteins, and lipids.
6. Oxidation/reduction
7. Acid/base properties and reactions
8. Equilibrium
9. Stoichiometry

RELATED SKILLS

1. Be able to set up, generate, and collect gases in the laboratory.
2. Be able to use common acid/base indicators.
3. Be able to balance molecular and ionic equations.

PERFORMANCE OBJECTIVES

After completing a study of biogeochemical cycles, students should be able to:

1. define biogeochemical cycle.
2. identify three major gaseous biogeochemical cycles.
3. distinguish between two major carbon "sinks"—carbon held in reserve in ocean sediments and in fossil fuels.
4. describe generally the movement of carbon, oxygen, and nitrogen in their respective biogeochemical cycles, between nonliving and living environments of earth.
5. relate how human intervention such as deforestation, the burning of fossil fuels, and industrial processes interrupt normal biogeochemical cycles.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Carbon Dioxide

Introduction

CO_2 plays a major role in the biogeochemical carbon cycle. You will examine some properties of CO_2 and its derivatives to develop an appreciation of the many roles this compound plays in the carbon cycle.

Purpose

To prepare carbon dioxide and study some physical and chemical properties of carbon dioxide and carbonates.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Immediately wash off any chemical with water if it contacts your skin.
3. Follow all normal precautions associated with the use of strong acids and bases.
4. Waft gases gently toward your nose when testing for odors.
5. Insure that proper care is used when lubricating and inserting glass tubing or thistle tubes into stoppers.
6. Make sure that the thistle tube is *below* the liquid level in the gas generator to prevent possible loss of CO_2 and spewing of acid.
7. Do *not* look directly at burning magnesium, since ultraviolet radiation, as well as the brightness, may cause eye injury.
8. Do not discard solids in the sink. Place in container provided by your teacher.

Procedure

1. Set up four test-tubes in a rack. Place approximately 2.5 g of calcium, magnesium and potassium carbonate in each of three tubes and sodium bicarbonate, $NaHCO_3$, better known as "baking soda," in the fourth. Add 3-4 mL of water to each and stir each mixture to effect maximum dissolution.
2. a. Obtain 75 mL limewater, a saturated solution of $Ca(OH)_2$, in a 250-mL beaker.
   b. Add a drop or two of 3.0 M hydrochloric acid, HCl, a strong acid, to each test-tube, fit rubber stopper into tube, and bubble any gas evolved into the limewater as shown in Figure 1. This is a standard qualitative test for carbon dioxide, CO_2.

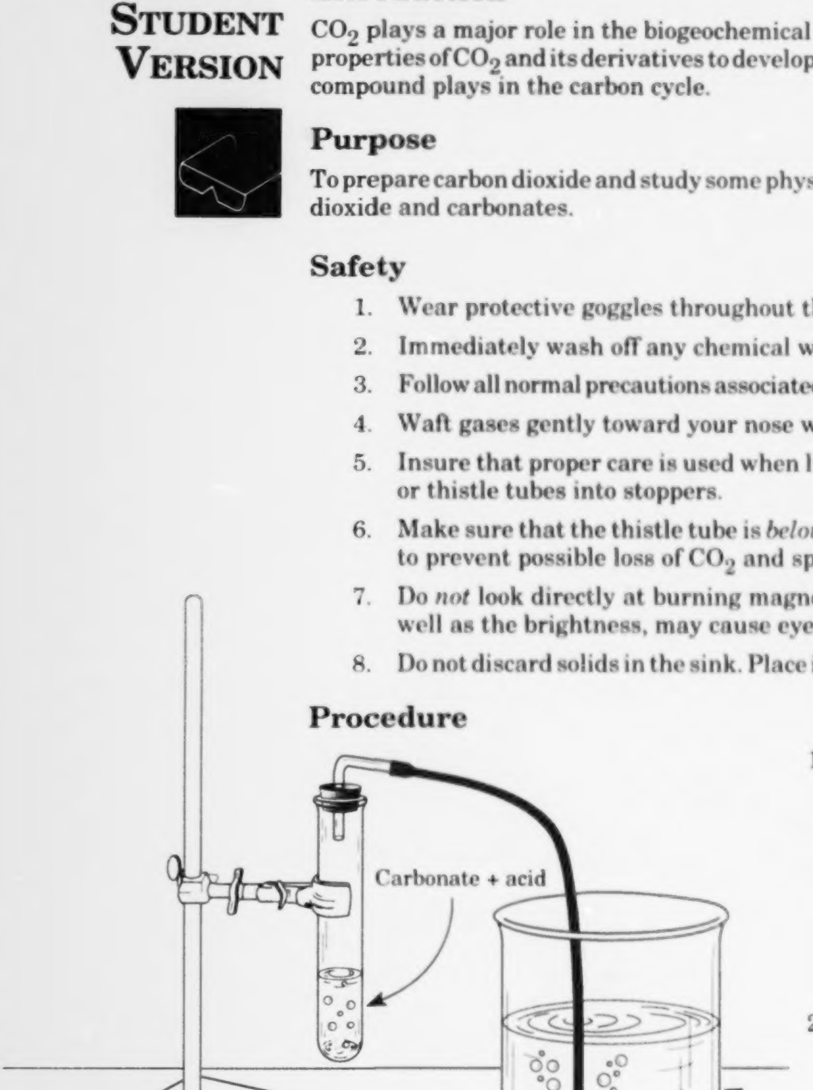


Figure 1. Bubbling $CO_2(g)$ into limewater.

3. Set up the CO_2 generator as shown in Figure 2. Before adding any acid, be sure to have your instructor check your set-up.

   Just cover the bottom of the gas generator bottle with marble chips, and carefully add enough (about 10 mL) 3.0 M HCl through the thistle tube to cover the marble. Be sure the thistle tube is below the acid level.

4. Discard the first bottle of gas collected. Adding 10 mL portions of 3.0 M HCl as needed, collect three bottles of CO_2, and confirm that you are indeed producing carbon dioxide by bubbling the gas into a test-tube of 3-4 mL limewater.

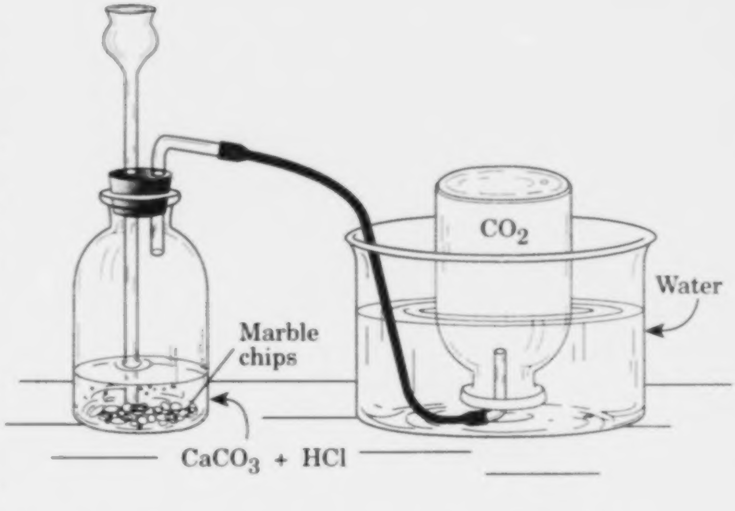


Figure 2. Laboratory preparation of carbon dioxide.

5. Continue to allow CO_2 to bubble into the tube of limewater after the initial confirmation that CO_2 is being produced. What happens now? What you observe occurs because calcium bicarbonate, $Ca(HCO_3)_2$, is much more soluble than calcium carbonate, $CaCO_3$. The clear solution is now "temporary" hard water.
6. Shut down your CO_2 generator by unstoppering and filling the generator bottle with water. Carefully decant this diluted acid, and rinse the marble chips several times with water. *No solids in the sinks.* Place the marble chips in the container provided by your teacher for reuse.
7. **Bottle 1**: Set up a 250-mL beaker containing a *short* burning candle inside. Being careful not to drip excess water, "pour" (as you would water) a bottle of CO_2 into the beaker over the flame. What do you observe?
8. **Bottle 2**: Holding a short piece of magnesium ribbon with small forceps, ignite in a burner flame, and *immediately* thrust the burning magnesium into a bottle of CO_2. *Do not drop the ribbon as it might break the bottle.* What do you observe?
9. **Bottle 3**: Into the last bottle of CO_2, add 50 mL of water and, covering the mouth tightly with a stopper, shake vigorously. What evidence of dissolving do you observe?
10. Make a *very* dilute solution of $NH_3(aq)$ (one drop 3.0 M $NH_3(aq)$ in approximately 400 mL water) and add 2-3 drops of phenolphthalein. The solution should be *very* pale pink. Now add your $CO_2(aq)$ solution to 25 mL of the basic solution. What happens?
11. Test aqueous solutions of $NaHCO_3$ and Na_2CO_3 with pH paper, litmus paper, and phenolphthalein. Describe the results.
12. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Which substances in Step 1 are soluble and insoluble?
2. Is a gas evolved when HCl is added in Step 2?

3. What happens when the gas in Step 2 is bubbled into limewater?
4. Bubbling CO_2 into limewater is a standard qualitative test for CO_2. State the test in your own words.
5. Write an equation for the reaction of CO_2 with limewater.
6. Why must the thistle tube be below the acid level in the CO_2 generator?
7. Why should you discard the first bottle of CO_2 collected in Step 4?
8. Write one equation for the reactions in Steps 4 and 5 using double arrows.
9. How does Le Chatelier's Principle apply to Steps 4 and 5?
10. How is the reaction from Question 8 a contributor to the hardness of natural water?
11. What physical and chemical properties of CO_2 make it a good fire extinguisher?
12. Explain your observations in Step 7 in terms of these properties.
13. What are the white powder and the black deposits on the sides of the bottle in Step 8?
14. How could these be formed?
15. Write a balanced equation for this oxidation-reduction reaction.
16. Describe any evidence of dissolving observed in Step 9.
17. What color change occurred in the $NH_3(aq)$ solution when you added the $CO_2(aq)$ in Step 10?
18. What does this color change indicate about a solution of CO_2?
19. Describe the results of your tests in Step 11.
20. Is the hydronium ion concentration of Na_2CO_3 greater or smaller than that of water? Is the hydronium ion concentration of $NaHCO_3$ greater or smaller than that of water?
21. Which compound has the higher pH, Na_2CO_3 or $NaHCO_3$?
22. Explain your answer to Question 21 in terms of hydrolysis and Le Chatelier's principle.

Implications and Applications

1. Discuss why CO_2 as a fire extinguishing substance would be appropriate or inappropriate in each of the following:
    a. An alcohol fire in a chemistry laboratory.
    b. A thermite reaction (burning aluminum and magnesium) gone out of control in the laboratory.
    c. An electrical fire in the roof of an elevator containing people stuck between floors.
    d. A grease fire in your kitchen.
2. Discuss the baking alternatives: baking soda and yeast. Is there any advantage in using one or the other?

3. Define hard water, distinguishing between the "permanent" and "temporary" kinds. "Temporary" hard water, chiefly composed of calcium and magnesium bicarbonates, can be "softened" by boiling or by the addition of sodium carbonate. Discuss the mechanism in each method that removes the "hard" ions. Which type would you prefer in your community and why?

4. In emergency rooms, a shout of "Bicarb!" is often heard, especially for victims of cardiac arrest or strangulation. What reason would there be for intravenous "bicarb" infusion?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Carbon Dioxide

Major Chemical Concept

Students should develop an appreciation for the many roles CO_2 plays in the biogeochemical carbon cycle—acid former, important constituent in carbonate-bicarbonate equilibria, inorganic carbon source in photosynthesis, and respiration product. Additionally, by observing carbon dioxide's density, nonflammability, and inability to support combustion, they will more clearly understand its uses in fire fighting, as well as some of its disadvantages and hazards under certain conditions (such as metal fire). Students apply previously learned concepts in new and practical situations: acid-base theory, hydrolysis of salts, Le Chatelier's Principle and equilibrium, oxidation-reduction, and writing equations.

Level

This activity is appropriate for any general first-year chemistry class.

Expected Student Background

Since this is an enrichment module, the material will probably be used late in the school year. It presupposes a wide range of conceptual understanding. Students should be familiar with Brønsted-Lowry acid-base theory, hydrolysis, concentration effects in equilibrium (Le Chatelier's Principle), oxidation-reduction, and how to balance equations. In addition, they should have a practical knowledge of laboratory techniques, specifically the generation and collection of gases.

Time

This activity can be completed in two 50-min periods, but it should be done over three periods if you use extended discussion to show applications and reinforce previously acquired concepts.

Safety

Read the *Safety Considerations* in the *Student Version*. Marble chips can be "recycled," preventing the dumping of solids in the sink as well as stressing conservation of resources and economy. Make sure the chips are rinsed well and then placed on paper towels to dry before putting back into container. Plastic thistle tubes are recommended instead of glass. Students should take care when lighting the short candle in the beaker. Use a splint, and do it quickly. Warn students not to look directly at burning magnesium, since it emits ultraviolet radiation that may harm their eyes. Also remind them not to drop the burning ribbon since it may crack the beaker.

Materials (For 24 students working in pairs)

Nonconsumables

12 Thistle tubes
48 Wide-mouth bottles
12 Pneumatic troughs
12 Graduated cylinders, 10-mL
12 Glass bends
12 Stoppers, 2-hole (sized to fit bottle)
12 Flexible rubber tubing pieces, 45 cm
48 Test-tubes, 25- x 150-mm
12 Test-tube racks

12 Spatulas
12 Beakers, 250-mL
12 Beakers, 400-mL
12 Forceps, small
36 Glass plates, 3" x 3", or large enough to cover the mouths of the collecting bottles
12 Rubber stoppers, solid, for wide mouth bottles
12 Rubber stoppers, 1-hole fitted with glass bends and rubber tubing No. 2
12 Stirring rods

Consumables

Sodium bicarbonate, $NaHCO_3$, 30.0 g
The following carbonates are suggested for Step 1. Check safety before using other carbonates:
  Potassium carbonate, K_2CO_3, 30.0 g
  Magnesium carbonate, $MgCO_3$, 30.0 g
  Calcium carbonate, $CaCO_3$, 30.0 g
Marble chips
12 Candles, small (short enough so they do not extend above top of 250-mL beaker)—birthday cake candles in clay blob "candle holders" work well
12 Splints
12 Magnesium ribbons, approximately 2 cm long
Litmus paper
pH paper
Phenolphthalein solution, 100 mL (1% in ethanol, divided into several dropping bottles)
3.0 M Hydrochloric acid, HCl, 1.0 L (250 mL 12 M HCl per 1 L solution)
3.0 M Ammonia, NH_3, 100 mL (20 mL 15 M NH_3 per 100 mL solution)
0.25 M Sodium carbonate, Na_2CO_3, 100 mL (2.7 g anhydrous Na_2CO_3 per 100 mL solution)
0.50 M Sodium bicarbonate, $NaHCO_3$, 100 mL (4.2 g anhydrous $NaHCO_3$ per 100 mL solution)
Saturated solution of limewater (Place enough $Ca(OH)_2$ powder or flakes—approximately 25 g—in a 1 L bottle deionized water to make the saturated stock solution. *Prepare 24 hours ahead.*)

Advance Preparation

1. Prepare solutions needed for the activity.
2. Check to be sure the $NH_3(aq)$-phenolphthalein solution students will prepare is dilute enough to be decolorized by the weak carbonic acid solution generated. If not, dilute the stock solution of NH_3 until you obtain the desired result.

Pre-Laboratory Discussion

1. You may wish to discuss CO_2 in these terms prior to the laboratory activity. CO_2 plays a major role in the biogeochemical carbon cycle. On the *geo* side its capture from the atmosphere accounts for most of the "sink" of insoluble mineral carbonates in the earth's crust and ocean floor. It also plays a major role in the production of "hard" water and helps buffer the oceans, maintaining a fairly constant pH for seawater through the production of bicarbonate ions. On the *bio* side it is essential to plants as the source of inorganic carbon from which organic carbohydrates can be synthesized, storing the energy of sunlight as chemical energy. CO_2 returns to the atmosphere as a by-product

of cellular respiration of carbohydrates and the release of their energy, but not before providing some important buffering activities in the circulatory systems of many organisms. Finally, CO_2 absorbs reflected infrared radiation from the earth's surface, targeting it as the major cause of global warming and the greenhouse effect. As the product of the burning of vast quantities of carbon-containing fossil fuels (along with water, also a greenhouse gas), the atmospheric concentration of CO_2 has steadily increased during the last two centuries. Whether or not natural geochemical processes can adapt to keep up with this increased atmospheric loading is debatable.

2. As part of this topic, stress concern for conservation of resources by example. "Recycle" the marble chips by rinsing and saving. This may seem trivial, but it does make an impact on students. Make sure the chips are rinsed well and then placed on paper towels to dry before putting back into container.
3. Caution students not to leave the top off of the limewater container. Why? *[CO_2 from the air will react.]*
4. Review safety emphasizing:
    a. Proper handling and disposal of acids.
    b. Proper set-up for gas generation and approval policy. (It should be your policy to inspect generators before the acid is added.)
    c. Warning about eye injury from looking directly at burning magnesium strip.
    d. Warning about lighting the short candle with care, using a splint.
5. To prevent gas from bubbling out through the thistle tube, emphasize that the tube should almost touch the bottom of the generator bottle so the liquid level completely submerges it. Sometimes an airlock prevents additional acid from going down into the generator. Tapping the bottle on the counter top gently, being careful not to slosh acid out, will break the airlock and allow the acid to flow. *Do not allow students to complete this procedure—only the teacher.* (If tapping does not work, briefly loosen the generator stopper; tighten immediately.)
6. Remind students to discard the first bottle of gas generated, as it will be contaminated with air. Three bottles of CO_2 can be collected easily in less than 15 min, although a recharge of acid may be necessary.
7. Several important concepts that have been studied previously should be reviewed here. Density (specific gravity) of gases, factors affecting equilibrium (Le Chatelier's Principle), Brønsted-Lowry acid-base concept (see respective *SourceBook* modules) will all play a role in understanding why and how some of the reactions to be observed work the way they do. You might also wish to review the meaning of acid anhydride, relating this concept to CO_2, the oxide of the nitrogen cycle and acid rain.
8. Emphasize the practicality of the knowledge to be learned and the many applications of these reactions.

Teacher-Student Interaction

First, walk around and check to be sure students understand how to set up the limewater test in Figure 1. More importantly, monitor insertion of glass tubing and thistle tubes into and out of stoppers. (Some teachers like to fit the rubber stoppers with glass tubing and thistle tubes in advance. If you choose to allow the students to prepare their own set-ups, be sure to show them the proper way of inserting glass tubing into a rubber stopper, including lubrication of the tubing with H_2O and use of a laboratory towel for hand protection.

An alternative way to fit the tubing is to insert a well-lubricated—with silicone grease— metal cork borer of larger diameter then the glass tubing into the stopper, slip the glass tubing through the borer and remove the borer, leaving the glass tube in place in the stopper. Although this method requires a little more advance planning, it is much safer than the standard methods.) Advise students concerning proper placement of the thistle tube (below liquid level), and remind them to discard the first bottle of gas collected. The reaction, once you have approved each set-up and students have added the acid, is rather rapid. As the rate of bubbling slows, instruct students to add more acid in 10 mL portions. Remind students to test the gas from the generator also, using limewater, and then stop the reaction. Walk around and make sure they do not leave marble chips reacting in the unstoppered bottles, do not dump solids into the sinks, or forget to rinse the marble and then put it on paper towels to dry.

As students begin to use their collected bottles of CO_2, monitor their work by walking around and asking questions: "Why did the candle go out?" "Why could you pour the CO_2?" "What does a higher pH mean?" Or make comments: "Be careful not to drop the burning strip." "That ammonia solution is too pink; dilute it more."

Anticipated Student Results

Many of the observations will be given in question responses to *Data Analysis and Concept Development* (see *Answers to Data Analysis and Concept Development*). In Step 2, a gas is evolved. This gas is suspected to be CO_2. At the end of Step 4, CO_2 is bubbled through limewater to produce a cloudy suspension as confirmation that the gas generated is CO_2. However, in Step 5, the cloudiness disappears (see *Post-Laboratory Discussion* for relevant equations). In Step 7, pouring CO_2 over the flame should extinguish the flame. In Step 8, a black solid (C) and a white solid (MgO) are formed when Mg is burned in CO_2. In Step 9, suction may be felt as the CO_2 dissolves in water. (Pressure decreases as gas dissolves.)

Answers to Data Analysis and Concept Development

1. **Soluble**: Na_2CO_3, $NaHCO_3$, $(NH_4)_2CO_3$

   Insoluble: $BaCO_3$, $CaCO_3$ ($MgCO_3$ is only slightly soluble.)
2. Yes
3. Limewater turns cloudy.
4. If an unknown odorless, colorless gas is bubbled into limewater and the clear solution becomes cloudy, the gas is CO_2.
5. $CO_2(g) + Ca(OH)_2(aq) \rightarrow CaCO_3(s) + H_2O(l)$
6. Otherwise, CO_2 bubbles out through the thistle tube instead of going into the collection bottles.
7. The first bottle is contaminated with air that was in the generator. Since it is contaminated with air, you might get false observations.
8. $Ca(OH)_2(aq) + 2\,CO_2(g) \rightleftharpoons CaCO_3(s) + CO_2(g) + H_2O(l) \rightleftharpoons Ca(HCO_3)_2(aq)$

   $[CaCO_3(s) + CO_2(g) + H_2O(l) \rightleftharpoons Ca(HCO_3)_2(aq) \rightleftharpoons Ca^{2+}(aq) + 2\,HCO_3^-(aq)]$
9. As the concentration of CO_2 increases, it drives the reaction beyond $CaCO_3$ formation to formation of $Ca(HCO_3)_2$, which is soluble.
10. Hard water contains $Ca^{2+}(aq)$, and soluble $Ca(HCO_3)_2$ adds $Ca^{2+}(aq)$ to water.

11. Physical property: density; denser than air. Chemical property: does not support combustion.
12. The denser CO_2 settles in the beaker, driving air out. This "smothers" the flame.
13. The white powder is MgO; the black specks are carbon.
14. To continue burning and forming MgO, magnesium literally tears CO_2 apart to remove the oxygen, leaving the carbon behind.
15. $2Mg(s) + CO_2(g) \rightarrow 2MgO(s) + C(s)$
16. Suction is felt on the palm.
17. The solution became colorless.
18. It has acidic properties; it neutralizes the base.
19. Litmus turns blue in both solutions, and phenolphthalein turns pink in both. pH paper indicates a higher pH for Na_2CO_3 than $NaHCO_3$.
20. The hydronium ion concentration of both Na_2CO_3 and $NaHCO_3$ is smaller than that of H_2O.
21. Na_2CO_3 has the higher pH.
22. CO_3^{2-} is a stronger base than HCO_3^-; the equilibrium representing CO_3^{2-} reacting with H_2O goes farther to the right. The pertinent equations are:

$$HCO_3^{2-}(aq) + H_2O(l) \rightleftharpoons CO_2(g) + H_2O(l) + OH^-(aq)$$

$$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^-(aq) + OH^-(aq)$$

Answers to Implications and Applications

1. a. A CO_2 fire extinguisher would be appropriate for use in a laboratory alcohol fire because it would smother the fire by cutting off the oxygen needed to feed combustion.

   b. A CO_2 fire extinguisher would not be appropriate with a thermite reaction since burning aluminum and magnesium would be hot enough to react with the CO_2 in a reaction similar to the one observed in the Bottle 2 (Step 8) reaction. Covering the burning mixture with sand is the appropriate method of removing oxygen here.

   c. A CO_2 fire extinguisher would not be appropriate in this situation since the density of carbon dioxide is so high. It would tend to settle to the floor of the elevator, displacing air, and creating a suffocating pocket of CO_2 that might harm passengers. This would be similar to a common situation faced by miners in coal mines before ventilation requirements became so strict.

   d. A CO_2 fire extinguisher would be appropriate in a kitchen grease fire in order to suffocate the combustion process. Water or a soda-acid fire extinguisher would be inappropriate, since the immiscibility of grease and water would result more in spreading the fire than cooling it.

2. Baking soda, $NaHCO_3$, dissociates in aqueous medium to $Na^+(aq)$ and $HCO_3^-(aq)$. The $HCO_3^-(aq)$ reacts with $H^+(aq)$ in the mixture to produce $H_2CO_3(aq)$ that, upon heating, produces $CO_2(g)$ and $H_2O(l)$. As the CO_2 diffuses through the dough, it creates bubbles that cause the dough to rise. After the CO_2 and water vapor leave the dough, the only residue is a sodium salt, which may not be desirable for people with high blood pressure. Note, also, that heating is necessary to cause the dough to rise.

In the case of yeast, enzymes (see *Enzymes* module) accelerate the metabolism of sugar according to the following reaction:

$$C_6H_{12}O_6(s) + \text{Yeast} \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

rendering heat unnecessary above normal temperatures. After the dough has risen because of diffusion of CO_2, baking stops the fermentation process by killing the yeast and evaporating off the alcohol. (Have you ever eaten rolls that tasted of alcohol? What might have caused this?) The lack of sodium and the rising of the dough without heat can be considered advantages of using yeast instead of baking soda. The disadvantage is the increased time that it takes to produce the bread.

3. Hard water contains dissolved minerals, usually containing calcium, magnesium or aluminum ions, but no bicarbonate anions. It does not lather easily and leaves a soap scum on tile. This scum is the insoluble "salt" of the soap. An example would be the conversion of soluble sodium laurate, a common soap found in some shampoos, into insoluble calcium laurate, which precipitates as a "soap film."

   In the case of "temporary" hard water, boiling decomposes the soluble bicarbonate into the insoluble carbonate, which removes the "hard"ions, leaving "soft" water. Replacement of bicarbonate with carbonate ions through addition of sodium carbonate accomplishes the same result. The fact that heating causes a precipitate to form makes temporary hard water a nuisance since insoluble carbonates build up in such household appliances as tea kettles, washing machine inlet pipes, and shower heads, and coffee makers.

4. In cardiac arrest or strangulation, the cessation of breathing causes a build up of CO_2 as the living cells continue to respire. This "dumping" of CO_2 into the bloodstream causes a drop in blood pH that itself can prove fatal:

   $$CO_2(g) + 2H_2O(l) \rightleftharpoons H_3O^+(aq) + HCO_3^-(aq)$$

   At first the blood's buffering capacity, which is due to blood bicarbonate concentration, attempts to maintain a constant pH by removing hydronium ions produced as CO_2 concentration increases by the reverse process:

   $$HCO_3^-(aq) + H_3O^+(aq) \rightleftharpoons CO_2(g) + 2H_2O(l)$$

   As the blood's bicarbonate concentration becomes depleted by this process, an infusion of "bicarb" is needed to try to maintain this buffering effect in the bloodstream.

Post-Laboratory Discussion

The discussion questions in *Implications and Applications* can lead to misconceptions or problems arising from the activity. You will particularly want to go through all reaction equations explaining what happened. (The limewater test reactions are given in Question 8 in *Answers to Data Analysis and Concept Development*. Hydrolysis reactions are in Question 23 in *Answers to Data Analysis and Concept Development.)* This is particularly important since equilibrium shifts and dissociation of weak and strong acids in the hydrolysis reactions may not be clearly understood. A possible extension of the post-laboratory discussion can center around CO_2 and its role in the "greenhouse effect."

Assessing Laboratory Learning

Besides the normal laboratory report, including answers to the questions asked, the following laboratory practical can be done.

A long candle is set up in a large beaker to which limewater has been added. A bottle of air is lowered quickly so that the mouth of the bottle is below the surface of the limewater, as shown in Figure 3. Account for the following observations in light of the CO_2 activity just completed:

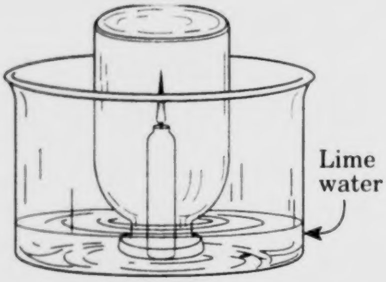


Figure 3. Burning a candle over limewater.

1. The candle slowly goes out. *[O_2, which is needed for combustion, is the limiting reactant. Burning ceases when it is depleted.]*
2. The limewater level rises inside the bottle. *[A combination of reaction of CO_2 with limewater and rapid cooling of hot gases, with the resultant decrease of volume, reduce gas pressure inside the bottle. Atmospheric pressure is now greater and forces limewater up into the bottle.]*
3. The limewater inside the bottle becomes cloudy at the surface only. *[This observation proves the CO_2 is produced by the burning of the candle, and that some of it is being removed by the production of insoluble $CaCO_3$. NOTE: The small amount of CO_2 in the original air could not account for the surface "film."]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Oxygen in the Air

Purpose

To measure the volume percentage of oxygen in moist air and calculate from it the percentage of oxygen in dry air.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. All chemicals should be treated with respect—even the most common are not exceptions. Be sure to wash off any chemical with water as soon as it contacts your skin.
3. Avoid breathing acetone vapors or spilling it on your skin.
4. Take care not to cut yourself while handling steel wool.
5. Washing of steel wool should be done only in the fume hood, using forceps.
6. Acetone is flammable; keep it away from flames.

Procedure

1. To each of two 20- x 150-mm test-tubes, attach lengthwise a strip of masking tape on which to record water levels. Set up a ringstand and two utility clamps, so that the inverted tubes can be set up in the 1000-mL beaker as shown in Figure 4. The beaker should be 2/3 full of water.
2. Weigh two 1.00 g pieces of size 00 steel wool, taking care not to compact the material.
3. *In the hood*, rinse in acetone (for about 30 s) one piece of steel wool, using small forceps. This removes any oil from the surface. Shake off any excess and drain on a paper towel in the hood. Then place the steel wool in the 25.0 mL of 1 M acetic acid provided in the hood and agitate for about 1 min. Shake off and drain as before on another paper towel.
4. Again rinse the steel wool, this time using 25.0 mL of 0.1 M acetic acid. (Agitate for 30 s, shake off excess and drain again on a fresh paper towel.)
5. Insert the cleaned steel wool into a test-tube, pushing it to the bottom with a stirring rod without compressing it. It should spread over the bottom half of the tube as shown in Figure 4. Immediately invert the test-tube and carefully lower it into the beaker of water. Tilt the tube and beaker so that approximately 2.0 cm of water goes up into the tube. Adjust the levels of water until they are the same inside and outside the tube, and clamp the test-tube in a vertical position.

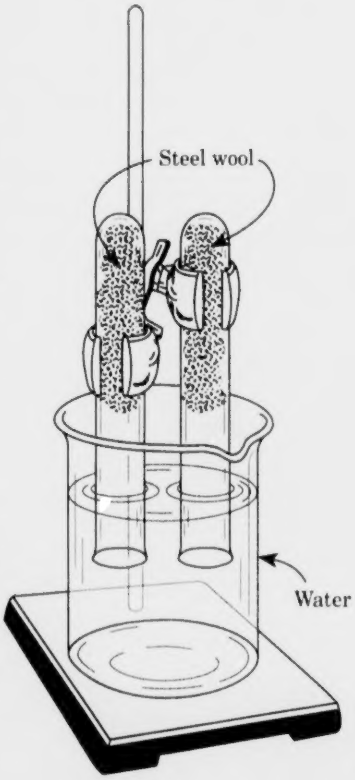
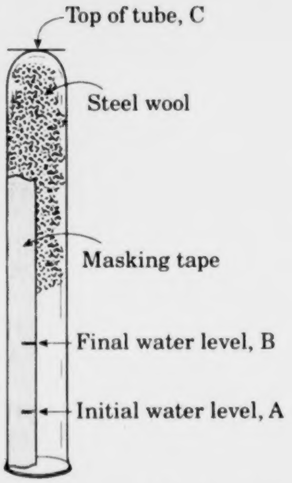


Figure 4. Apparatus for measuring the oxygen in air.



Original Amount of Gas in Tube [C-A]
Final Amount of Gas in Tube [C-B]

Figure 5. Water levels.

6. Mark the initial water level on your tape with a waterproof pencil or pen (A in Figure 5).
7. Repeat the entire procedure (Steps 3-6) for the other tube, taking care not to disturb Tube 1 as you set up Tube 2. The mouth of the tubes should be below water level throughout the activity.
8. When no further change in water level can be detected, after about 20-30 min (it is acceptable to leave the set-up overnight at this point), wait 5 min longer and then, by lowering the tube, readjust the water levels inside and outside until they are the same.
9. Mark the water level carefully on your tape (B in Figure 5), take an accurate temperature reading of the water in the beaker without touching the sides (we will assume the gas inside the tube and the water in the beaker are the same temperature), and a barometric pressure reading. Construct a data table in your laboratory notebook and record your data.
10. Drain the first test-tube, dry it, and accurately estimate the distance between the two lines to the nearest hundredth of a centimeter. Next, estimate the length from the initial water level to the bottom of the tube (A-C distance in Figure 5) to the nearest hundredth of a centimeter. Using these data calculate the percent by volume of O_2 in moist air.
11. Correct your percent value to that of dry O_2 using the following equation:

$$\text{Volume \% dry } O_2 = \text{Volume \% moist } O_2 \times \frac{P_b}{P_b - P_w}$$

where P_b is the barometric pressure, and P_w is the water vapor pressure at the temperature recorded.

12. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Why are two test-tubes used in Step 1?
2. Why must you be careful not to pack the steel wool in Step 2?
3. Why is it necessary to rinse the steel wool in acetone and acetic acid?
4. What is the purpose of equalizing water levels in Step 5?
5. Why did water rise into the tubes in Step 8?
6. Why is it necessary to adjust water levels again in Step 8?
7. Why do you need to know gas temperature and barometric pressure?
8. What equation did you use for the calculation of percent by volume of O_2 in moist air in Step 10?
9. Why is the air moist?
10. How does your corrected percentage of dry oxygen in air calculated in Step 11 compare with the accepted value of 20.9% volume O_2 for dry air at sea level? Calculate your percentage error.

Implications and Applications

1. In geology, iron deposits can be used to hypothesize atmospheric change and the evolution of life producing that change by the state of oxidation of the iron. For instance, in what order would you expect to find banded layers of FeO and Fe_2O_3 in deep sediments as you dig into the crust? What does this imply?
2. Ozone (O_3), an allotropic form of oxygen, plays a major role in screening harmful ultraviolet light from sunlight in the atmosphere. What implication is there for the change in banded iron deposits and the development of the ozone layer around the earth?
3. If you were to test the remaining gas in each of the test-tubes used in this activity with a burning splint, what would occur? Why? What gas would this be if all but a very tiny fraction reacted with magnesium to form Mg_3N_2?
4. Based on conclusions in Question 3, what can you imply about combustion and/or respiration in the atmosphere compared to that in a pure oxygen environment? Is it not fortuitous (a "lucky accident") that 78% of the atmosphere is composed of this "remaining gas"? Give an example to illustrate what might happen if it were not.
5. How would you design an activity to measure the percentage loss of O_2 and percentage gain of CO_2 in respired air?
6. Slow oxidation (rusting) of iron is a major economic problem. A person who found a "better way" to prevent or slow the process would probably become a billionaire overnight. Since this is an oxidation-reduction process, what are three methods currently used to retard the process?

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Oxygen in the Air

Major Chemical Concept

Earth's atmosphere is very different from the planet's original one composed chiefly of methane (CH_4) and ammonia (NH_3), with oxygen estimated to have been only 10^{-12} what it is today. Although the earliest oxygen produced seems to have been used to oxidize reduced atmospheric gases like CH_4, NH_3, H_2S, *etc.* and crustal materials, like the production of banded iron formations from iron rising from the core, photoautotrophs gradually increased the percentage of O_2 in the atmosphere over millennia. Today, there is an approximate "reservoir" of 10^{15} tons of atmospheric oxygen. Theoretically, all of the O_2 produced by plants should have been used by them in respiration. However, since some plant material produced became geologically "encapsulated" in the crust as fossil fuels and other organic substances, a surplus of oxygen gradually built up in the atmosphere.

The student should gain an insight into the important role O_2 plays and has played in the development of life on this planet. In addition the reactive nature of the element should be demonstrated, while a reasonably accurate quantitative determination of O_2 in today's atmosphere can be made using simple techniques. Also, a related discussion of ozone and its importance can be brought in as a related topic.

Level

This activity is appropriate for any general first year chemistry class.

Expected Student Background

Students should have studied the preparation and properties of oxygen and the gas laws. Particularly, they should know Dalton's Law of Partial Pressures and how to adjust for water vapor pressure. They will probably have studied the atmosphere in earth science, but this will be an effective way to verify previously learned facts.

Time

This activity can be done in one 50-min period, but it will be most effective if several hours are allowed for the reaction. Other strategies are to have students set up the activity after school the day *before* the activity, or set up in the laboratory and come back after school to take final readings and clean up. Set up takes about 20 min. Discussion after the activity may require another period.

Safety

Read the *Safety Considerations* in the *Student Version.*

1. Be sure the washing of the steel wool in acetone is done in a fume hood.
2. Students should avoid breathing of acetone vapors or contact with skin, and no burners should be in use nearby, since acetone is flammable.
3. Beakers containing 100 mL of acetone and 1.0 M acetic acid should be set up in the hood for washing the steel wool. You should change these after use by two or three students. A reagent bottle of 0.10 M acetic acid should be available for students to obtain the 25.0 mL needed for their final rinse for both pieces of steel wool. Dispose of used acetone properly.
4. Students should be warned that "tearing" pieces off of steel wool can shred the skin painfully.
5. Acetone should be allowed to evaporate in the hood.

Materials (For 24 students working in pairs)

Nonconsumables

12 Beakers, 100-mL
12 Beakers, 1000-mL
24 Test-tubes, 20- x 150-mm [Note: Using 50-mL burets (if available) in place of test-tubes will improve volume measurement.]
24 Utility clamps
12 Ringstands
12 Rubber stoppers, #2 (optional)
12 Waterproof pencils or pens
12 Stirring rods

Consumables

Acetone, 1 L
1.0 M Acetic acid, CH_3COOH, 1 L (57 mL of 17.5 M [glacial] acetic acid diluted to 1.0 L). *REMEMBER*: Add acid to water.
0.10 M Acetic acid, CH_3COOH, 1 L (5.7 mL of 17.5 M acetic acid diluted to 1.0 L)
Fine steel wool (size 00), 1 pkg
Wood splints
Masking tape

Advance Preparation

Prepare needed acetic acid solutions and secure other materials for the activity.

Pre-Laboratory Discussion

Discuss safety with students. Emphasize the importance of not compressing the steel wool and accurate marking of tape with a waterproof pencil or pen. A small amount of water should be in the test-tube when it is set up as a starting point. Tilting the tube in the beaker allows a few milliliters of water to replace air in the tube. Care should be taken not to remove the test-tubes from the beaker until all water levels have been adjusted and accurate markings made. Also mention that the reason water levels must be equalized is that the total gas pressure inside the tube is the same as the student reads on the barometer. Inform students where to find a table of equilibrium water vapor pressure values. If students are to do the optional splint test for oxygen, remind them to stopper tubes *before* removal from the beaker. Finally, waste steel wool should be disposed of properly, not in the sinks, and acetone should be evaporated in the hood. Warn students that there may be some rust staining in their tubes, which can be removed with oxalic acid solution.

Teacher-Student Interaction

Students should be monitored at the hood to insure proper washing of steel wool and handling of acetone. Walk around and make sure students allow a small amount of water into the tube *before* clamping. Also insure that tubes are in a vertical position and that the starting position is marked after set up.

Anticipated Student Results

Students should be able to obtain a percent dry oxygen in air within 10% of the accepted value (20.9%).

Answers to Data Analysis and Concept Development

1. Duplicate set-up in case of misfortune.
2. Since O_2 reacts at the surface of the steel wool, and since it is important that all O_2 in the tube react, we do not want to limit the amount of iron exposed to the air in the tube by compressing the steel wool.
3. To insure that the steel wool is clean and free of oil and corrosion. Acetone will dissolve oil. Acetic acid will insure a clean surface.
4. Equalization is done to insure that the gas pressure inside the tube is known and, therefore, is the same as the barometric pressure.
5. As O_2 was removed from the air, the total gas pressure in the tube decreased, allowing the greater air pressure to force water into the tube until pressures inside and outside equilibrated.
6. Same answer as for Question 4.
7. The temperature is necessary to obtain the equilibrium water vapor pressure in order to adjust total pressure. The barometric pressure is equal to the total pressure inside the tube when the water levels are the same.
8. $$\text{Volume \% moist } O_2 = \frac{\text{Length water level rose}}{\text{Total length of original air column}} \times 100\%$$
9. The air is moist because it was over water, and some water vapor was present.
10. Answers will vary, and they should be less than 10% error.

Answers to Implications and Applications

1. Closer to the surface you would expect to find Fe_2O_3, with FeO deposits buried underneath these layers. This implies that the early atmosphere contained less oxygen and there was only enough to oxidize the crustal iron to the 2+ state, while later sediments indicate a greater abundance of oxygen, allowing oxidation to the higher 3+ state.
2. The increase in oxygen in the atmosphere indicated by the banded iron deposits would imply a concurrent increase in ozone production, resulting in the formation of a significant ozone layer around the earth:

   $$3O_2(g) \rightleftharpoons 2O_3(g)$$

   Since this layer would increasingly screen harmful ultraviolet wavelengths from sunlight, life could begin to move out from the ocean depths into the direct sunlight on land.
3. The burning splint would immediately go out, indicating that the gas(es) remaining after all of the oxygen has been removed do not burn or support combustion (respiration). The formation of magnesium nitride indicates that most of the remaining gas was nitrogen.
4. An object should burn much faster in pure oxygen than in air, since air is only approximately 20% oxygen (see *Reaction Rates* module). If each lungful of "air" were 100% oxygen, we would literally hyperventilate ourselves to death, since the cells would probably not be able to withstand the temperatures produced or resist being oxidized themselves by rapid reaction.

5. **For O_2**: Using the set-up in *Activity 2*, substitute a tube of respired air and perform the experiment using the same procedures, making sure that temperature and pressure conditions are the same as when air alone was used. The average weight-gain difference divided by the weight-gain of the iron in air times 100% gives the percent loss of oxygen in respired air. (The design of an apparatus containing only respired air would require a two-hole stopper and inlet and outlet tubes to the test-tube.)

   For CO_2: Invert equal volumes of air and respired air at the same temperature and pressure conditions in a container of limewater. After standing, and the formation of a precipitate, note any difference in water levels in the two containers. This difference could be used to find the percent volume difference of CO_2 in the two samples by dividing the difference by the volume of CO_2 removed from air and multiplying by 100%. Anything over 100% indicates the percent increase of CO_2 in respired air.

   Obviously several trials to obtain averages and large enough samples to produce measurable results must be used.

6. Three methods are: (1) lubrication, to exclude oxygen; (2) painting, also to exclude oxygen; and (3) use of a more reactive (than iron) metal to act in a sacrificial capacity and undergo oxidation in the redox process of corrosion. Other methods that prevent oxidation of the iron may be mentioned.

Post-Laboratory Discussion

Use *Implications and Applications* as a basis for discussion. Ask students if this reaction would have occurred without water being present. Then ask students to balance the equation: $Fe(s) + O_2(g) + H_2O(l) \rightarrow Fe_2O_3 \cdot x\ H_2O(s)$. (The x moles of water appearing on both sides of the equation indicates that iron oxide hydrate can have a variable amount of water.) Discussion of the stoichiometry of photosynthesis and respiration is also appropriate here: why there should be no surplus of oxygen in the atmosphere, and what the fact that one exists implies. This fact can also be used to illustrate the living (bio-) and nonliving (geo-) components of the oxygen cycle, their interrelationships, and how important the living portion of the cycle (rain forests, phytoplankton, *etc.*) are to it. Ozone and its depletion is another way to point out possible consequences of human intervention in the natural cycles on earth. "Spaceship Earth" has many backup systems to try to compensate for "system breakdowns." What happens when these backups fail due to human activities can seriously impact natural systems. If time permits, possible sources of error might be discussed.

Assessing Laboratory Learning

Percentage volume of dry oxygen in air should be 20.9%. Since this activity is intended to give an approximate percent volume, a 10% error is acceptable. Students should be required to explain possible sources of error that might cause too high or too low readings.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Carbon as a Component of Sugar

Purpose

Carbon, oxygen, and hydrogen are combined in the presence of chlorophyll and sunlight in the plant to produce carbohydrates. In this demonstration, you will use sulfuric acid to dehydrate the sugar, leaving carbon.

Materials

Small beaker, 150-mL (It will be difficult to clean afterwards, so you might want to use an old one.)
Beaker, 400-mL
Sucrose, $C_{12}H_{22}O_{11}$ (granulated table sugar)
Sulfuric acid, H_2SO_4, concentrated, 10 mL
Beaker, 600-mL

Safety

This demonstration will generate heat, which would possibly crack the beaker, and copious noxious fumes consisting of water mixed with sulfuric acid. *It must be done in a hood.*

Handle concentrated sulfuric acid with extreme care.

Use gloves, safety goggles, and a safety shield.

Procedure

1. Fill the small beaker with "30 mL" sugar and place this beaker inside a 600-mL beaker containing about 100 mL H_2O.
2. Measure the temperature of the H_2O in the larger beaker.
3. Carefully pour sulfuric acid onto the sugar until there is a layer about 1/8 of an inch deep (about 10 mL). *Do not stir.*
4. Set the beaker system aside in the hood and watch for a reaction to occur as evidenced by a color change where the acid contacts the sugar. The whole mixture will "puff up" dramatically.
5. After the reaction is completed, measure the temperature of the water again. The highly exothermic reaction should increase the water temperature. *[10 mL conc H_2SO_4 produces about a 10 °C rise in temperature.]*

Remarks

This reaction is a dehydration reaction represented by the following equation:

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2SO_4\text{ (conc)}} 12C(s) + 11H_2O(g)$$

A secondary reaction explains why the carbon forms a solid foam in the beaker. The escaping gases cause the foaming.

$$C(s) + 2H_2SO_4(l) \rightarrow CO_2(g) + 2SO_2(g) + 2H_2O(g)$$

Demonstration 2: The Solubility of Carbon Dioxide and Ammonia

Purpose

A major part of carbon and oxygen cycles involves dissolution of carbon dioxide gas in aqueous solution. In fact, the huge amount of carbon dioxide dissolved in the oceans provides a valuable "sink," or reserve supply, of carbon dioxide. Ammonia, one of the more important compounds of nitrogen is also part of the nitrogen cycle. Two demonstrations are suggested here to show the extent of solubility of these two gases and the effect of temperature on their solubility.

A. Carbon Dioxide

Materials

Fresh bottle of soda pop, preferably chilled (10 oz bottle is sufficient)
Empty plastic soda bottle (10 oz)
One-hole rubber stopper (size to fit soda bottle)

Glass tubing
Rubber tubing, 60 cm
Pneumatic trough
Large beaker filled with warm water
Limewater, 25 mL (optional; see *Activity 1* for preparation)

Safety

No special safety precautions are required.

Procedure

1. Prepare a one-hole rubber stopper to fit the mouth of a 10 oz plastic bottle of soft drink, like "Sprite" or seltzer water. Put a glass tube in the stopper hole and connect a 60-cm piece of rubber tubing to it. Prepare a water trough to collect gas by water displacement. Fill an empty 10 oz soft drink bottle with water, place it inverted in the water trough. Position the tube through the mouth and into the bottle.
2. Open a fresh bottle of precooled soda pop, immediately insert the rubber stopper, and gently shake the bottle and warm it by placing it in a large beaker of warm water.
3. Collect the carbon dioxide by displacing the water in the inverted bottle. *(Students will be amazed at the amount of gas evolved and collected. You will be able to completely fill the second bottle with gas.)*

Extensions

To prove that this gas is indeed carbon dioxide, remove the bottle containing the gas and immediately add about 25 mL limewater. A cloudy white precipitate of calcium carbonate is a positive test for carbon dioxide.

$$CO_2\,(g) + Ca(OH)_2\,(aq) \rightarrow CaCO_3\,(s) + H_2O\,(l)$$

Remarks

At 20 °C, 0.9 L of CO_2 will dissolve in 1 L of water.

B. Ammonia (The Ammonia Fountain)

Purpose

To test the solubility of ammonia gas in hot and cold water.

Materials

Aqueous ammonia, NH_3, concentrated, 10 mL
1000-mL beaker of hot water
1000-mL beaker of cold water
Florence flask, 500-mL
One-hole rubber stopper to fit Florence flask
Burner
Glass tubing
Phenolphthalein solution, few drops (optional)
Vinegar, few drops (optional)

Safety

Students should avoid inhaling ammonia vapor and should not handle hot glassware with their hands.

Procedure

1. Prepare a one-hole rubber stopper with 20-cm piece of glass tubing to fit a large (500-mL) Florence flask.

2. Place about 10 mL concentrated aqueous ammonia in the flask and insert the stopper. Have a large beaker of hot water and a large beaker of cold water nearby (Figure 6).
3. Gently heat the flask with a burner for 10-15 s and immediately invert it and place the stoppered mouth into the beaker of hot water. *(Students will observe that no water from the beaker enters the flask.)*

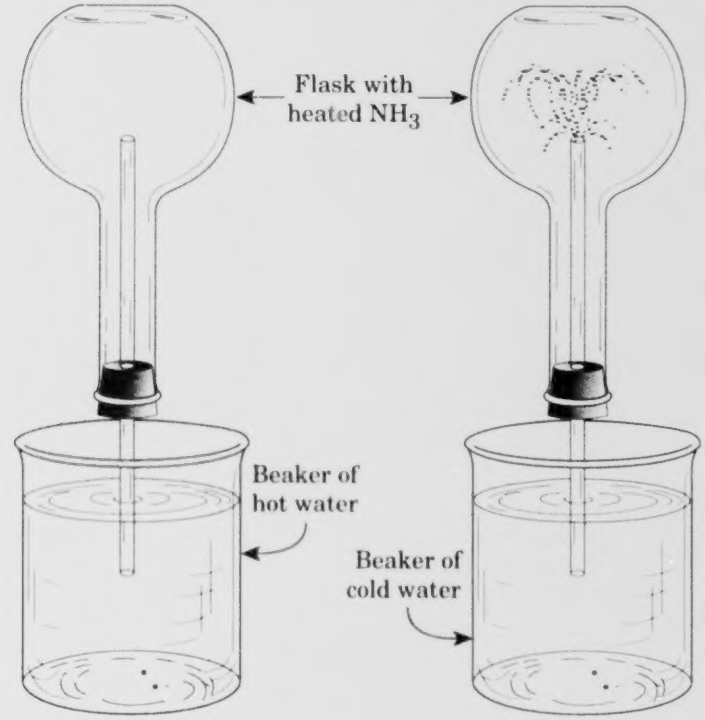


Figure 6. Ammonia fountain demonstration.

Remove the flask, heat it again for a few seconds, and invert it in the beaker of cold water. *(The water in the beaker will enter and fill the flask.)*

Extensions

You might want to add a few drops of phenolphthalein indicator to the water in the beakers and a few drops of vinegar. When the dilute acid containing phenolphthalein enters the flask, it will turn pink.

Remarks

Ammonia readily dissolves in cold water, but its solubility in warm water is less. Ammonia in solution is called "aqueous ammonia." It is incorrect to call this solution "ammonium hydroxide," since such a molecule has never been shown to exist (see *Acids and Bases* module).

Demonstration 3: Formation of NO_2 and N_2O_4 Gases

Description

This demonstration illustrates the $NO_2 - N_2O_4$ equilibrium.

Materials

Nitric acid, HNO_3, concentrated, 10 mL
Copper penny, pre-1982 or 1-2 g Cu wire
Side-arm flask, small
Rubber tubing
Two test-tubes
Rubber stoppers, 3 (one to fit the side arm flask, two for the test-tubes)
Beaker of boiling water
Beaker with ice bath
Strapping tape

Safety

Concentrated nitric acid is corrosive. You must wear gloves and a face shield when using this acid. NO_2 is a very toxic gas. It must be generated in a *hood*. The tubes should be stoppered tightly and sealed with tape. It has been estimated that as few as five pennies could generate enough NO_2 to reach a toxic level in a school laboratory.

Procedure

A. Preparing the Equilibrium Tubes

1. Prepare a gas generator by attaching a rubber tube to a small side-arm flask.
2. *Working in a hood*, add about 10 mL concentrated nitric acid to the flask. Drop in a pre-1982 penny. A reddish-brown gas, NO_2, will form immediately. Allow enough of this gas to be formed to displace the air in the flask and stopper the flask.
3. Fill two test-tubes with gas from the generator and immediately stopper the tubes tightly. Color intensity should be the same in both tubes initially.
4. Stop the reaction in the flask by filling it with water. (Note the blue color of the solution formed.)

B. Using the Equilibrium Tubes

1. Place one of the tubes in a beaker of boiling water. Notice the change in color of the gas in the tube.
2. Place the second test-tube in an ice bath. Notice the change in color of the gas in this tube.

Remarks

1. When copper reacts with concentrated nitric acid, reddish-brown, toxic $NO_2(g)$ is formed.

$$Cu(s) + 4H^+(aq) + 2NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_2(g) + 2H_2O(l)$$

The equilibrium mixture in the tubes consists of NO_2 and N_2O_4 gases, often referred to an NO_x. They react according to the equation:

$$2NO_2 \text{ (reddish-brown)} \rightleftharpoons N_2O_4 \text{ (colorless)} + \text{Heat}$$

When the mixture is heated, the equilibrium shifts toward the formation of brown NO_2. When it is cooled, equilibrium shifts toward the formation of colorless N_2O_4.

2. When dry NO_2 is collected, as in this demonstration, oxygen in the air maintains the nitrogen in the N(IV) state. When NO_2 is collected over water, a different set of reactions occurs (see *Industrial Inorganic Chemistry*, Demonstration 2).

Other Demonstration Ideas

Refer to the *Simple Chemical Reactions* module for additional demonstrations.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What is the difference between a dynamic equilibrium and a steady state equilibrium? *[A dynamic equilibrium occurs within a closed system only. It involves changes in which the same atoms move back and forth, and is affected as Le Chatelier's Principle predicts. A steady state equilibrium occurs in an open system only. It involves different atoms that enter and leave at the same rate with no net change, and does not usually behave as Le Chatelier's Principle would predict.]*
2. We have all heard of "nitrogen-fixation." Is it possible to "fix" or "convert to a more useful condensed state" other gases? *[Certain bacteria "fix" atmospheric nitrogen as more useful solid nitrates (usually in solution) in the roots of legumes. Photosynthesis, on the other hand, takes atmospheric carbon dioxide and "fixes" it as carbohydrates (usually in solution).]*
3. What source produces the greatest amount of carbon dioxide released to the atmosphere? *[Contrary to popular belief, respiration or combustion processes* do not *release the largest amount of CO_2 to the atmosphere. Aerobic fungal and bacterial decomposition of dead organic matter is the major source of atmospheric CO_2.]*
4. Compare the geologic processes that change limestone into marble with those that change peat into anthracite coal. *[Heat and pressure, created by deep burial or volcanic upheaval, literally melt sedimentary limestone to allow it to become metamorphic marble. Individual particles become fused into one mass. In coal formation, the same events create pressure and heat that, coupled with the initial action of aerobic and anaerobic bacteria, volatilize impurities and fuse the carbon particles together more and more over time. Bituminous, or soft coal, is "dirty" and rubs off easily on a white glove. Anthracite, or hard coal, that has been exposed to the greatest amount of metamorphism, is shiny and "clean," not producing much coal dust since the carbon particles have been melted and fused. That is why its carbon will not easily rub off on a white glove. Removal of volatile impurities gradually allows the coal to become purer carbon, forming a more perfect graphite ring structure.]*
5. Even if there were no industrial pollution, rainwater would still be weak "acid rain." Why? *[As rainwater falls through the atmosphere, CO_2 dissolves into it. Since CO_2 is an acid anhydride, it reacts with water to form weak carbonic acid, H_2CO_3. Small amounts of the stronger nitric acid, HNO_3, may be produced during lightning storms as N_2 is changed to NO and then to NO_2, that reacts with water to form nitric acid.]* (See *Industrial Inorganic Chemistry*, Demonstration 2.)
6. What role has photosynthesis played in creating the reservoir of oxygen in the earth's present atmosphere? *[O_2 is released by plants as they undergo photosynthesis. This oxygen, in turn, should be used during respiration to oxidize carbohydrates back to CO_2, which is then taken in by plants for photosynthesis (with a release of more O_2), completing the cycle. If carbon is trapped in an oxygen-poor environment, as it was in coal and petroleum formation, the oxygen released originally is not used in respiration, causing the reservoir (surplus) of O_2 to build up in the atmosphere.]*

7. Why is the "greenhouse effect" somewhat inaccurate in describing atmospheric warming as certain gases build up in the atmosphere? *[In a true greenhouse the heating effect is caused by the fact that a barrier (glass) will not allow the heat to be transferred by convection. In the atmosphere, there is absorption of certain infrared frequencies, which heat the atmosphere. There is no physical barrier preventing convection of the "heated atmosphere."*]

8. What role does CO_2 play in the transfer of insoluble limestone ($CaCO_3$) underground in order to form stalactites and stalagmites?

   A. $CO_2(g, \text{excess}) + H_2O(l) \rightarrow H_2CO_3(aq)$

   B. $H_2CO_3(aq) + CaCO_3(s) \rightarrow Ca(HCO_3)_2(aq)$

   C. $Ca(HCO_3)_2(aq) \rightarrow CaCO_3(s) + H_2O(l) + CO_2(aq)$

   [In Equation A, an excess of CO_2 dissolves in rainwater forming carbonic acid. This weak acid reacts with insoluble limestone above or in the ground to form soluble calcium bicarbonate or "temporary" hard water, shown in Equation B. The $Ca(HCO_3)_2(aq)$ moves through the soil in various ways, sometimes as droplets hanging from the ceiling of a cave. As conditions favor drying, the droplet evaporates, shifting the equilibrium and leaving behind a layer of limestone ($CaCO_3$) on the cave ceiling or floor and releasing carbon dioxide and water as shown in Equation C.]

9. "The nitrogen cycle is dependent at each step upon microorganisms and could not proceed without them." Support or refute this statement. *[Bacteria are necessary in each of the four steps in the nitrogen cycle. N_2 is initially "fixed" in the form of ammonia (NH_3) or ammonium compounds by soil bacteria, such as* Rhizobium japonicum, *or certain blue green algae. In Step 2 of the cycle, nitrification occurs, where chemosynthetic bacteria oxidize these ammonia based compounds into nitrites and nitrates. Plants use the latter to produce amino acids, which are used by animals to produce animal protein. At death, both plants and animals are decomposed by bacteria, releasing ammonia (ammonium compounds) to the soil. In the final step of the cycle, anaerobic bacteria break down nitrates and ammonium compounds to free nitrogen and the cycle begins again.]*

Counterintuitive Examples

1. The solubility of ammonia decreases (rather than increases) as the temperature of the water increases. Unlike solids dissolved in liquids, gases normally become more soluble in water as the temperature drops.

2. When burning a candle over limewater, the rise of limewater into the bottle is mistakenly attributed to the removal of O_2 from the air inside the bottle by the burning candle. In fact, limewater enters the bottle because of the removal of some CO_2 produced by the burning candle as it reacts with limewater and the contraction of CO_2 as it cools. This volume decrease creates less pressure inside the bottle, allowing atmospheric pressure to force additional limewater into the bottle.

Metaphors and Analogies

1. "Spaceship Earth." Like a spaceship, the earth is finite in size and capacity, has limited oxygen, water, and food supply aboard, and does not possess inexhaustible energy supply.

2. "Greenhouse Effect." In a greenhouse, glass transmits visible light and absorbs infrared radiation from the sun. This is changed to heat and prevented from leaving the greenhouse because it cannot penetrate the glass. Thus, the inside temperature of the greenhouse increases. (The same thing happens on a hot day when the windows of an automobile are up.) In the "atmospheric" greenhouse, there is no glass to absorb infrared rays, but there is a layer of several gases (primarily CO_2 and H_2O) that serve the same purpose. These gases trap heat and warm the earth without a physical barrier present.
3. Propose a familiar cycle (water in an automobile cooling system, *etc.*). Ask students to suggest things that might interfere with this cycle. Extend this analogy to cycles of carbon, oxygen, and nitrogen.
4. Consider the human circulatory system as similar to a biogeochemical cycle. What happens when an artery that serves a certain organ (heart, for example) is blocked? *[Some damage to the heart may result, and an alternate pathway develops due to collateral circulation.]* Compare this to events that might interfere with normal cycles of oxygen, carbon, and nitrogen.

TIPS FOR THE TEACHER

Background Notes

Introduction

Before discussing specific examples of biogeochemical cycles, *viz.*, the carbon, oxygen, and nitrogen cycles some terms must be defined: cycle, biogeochemical cycles, dynamic equilibrium, and steady state equilibrium.

Many agree with the dictionary definition of cycle as "a single complete execution of a periodically repeated phenomenon." **Biogeochemical cycles**, then, involve movement of elements back and forth between the physical (nonliving or abiotic) and living (biotic) components of the earth, considered to be a **closed system** (nothing enters or leaves the system) in most cases. Parts of the cycle function as a **dynamic equilibrium**, where reversible changes occur within a closed system as predicted by Le Chatelier's Principle (see *Chemical Equilibrium*). For example, the carbonate-bicarbonate interaction in the oceans serves to maintain seawater at a relatively constant pH. There is some movement of the same atoms back and forth within the balanced system. On the other hand, portions of the cycle function as a **steady state equilibrium**, not always functioning according to Le Chatelier's Principle, because there is only movement in one direction (irreversible change), and the system is considered to be open. A burning candle illustrates this definition. Since the number of carbon atoms from the wax (or oxygen atoms from the air) is the same as the number leaving in the CO_2 gas produced, we have a steady state equilibrium. The number of atoms of carbon (or oxygen) input equals the same number being output, so an equilibrium exists, though here they can never be the same atoms, since all CO_2 is lost to the atmosphere (an open system here). An understanding of the difference between dynamic and steady state equilibria is essential to understand the differing effects each input (or output) flux or change may have on a reservoir of an element.

Carbon Cycle

Carbon is one of few elements found in all three "spheres" of the earth's crust: lithosphere (solid), hydrosphere (liquid), and atmosphere (gas). In the lithosphere it can be found as coal, petroleum, limestone, and other carbonate rocks; in the hydrosphere in shells, coral reefs, and hydrogen carbonate ions dissolved in water; in the atmosphere primarily as CO_2.

The carbon cycle is the most important of the biogeochemical cycles; it plays the major role in energy transfer within the biosphere. **Photosynthesis**, primarily an energy storage mechanism for green plants, traps carbon dioxide from the atmosphere and stores energy from the sun in the form of carbohydrates. In essence, CO_2 becomes **fixed**, or transformed into a more useful, solid form (carbohydrates) by the action of living things:

$$6CO_2(g) + 6H_2O(g) \rightarrow C_6H_{12}O_6(s) + 6O_2(g)$$

Carbohydrates can be converted into proteins, fats, and other complex organic compounds.

Animals, unable to perform photosynthesis, must depend on green plants to provide their food, incorporating the carbon components from food into many different carbon compounds in the animal body. In the same way, carbon is transferred and reused by carnivores who eat herbivores or omnivores.

Life would cease if all carbon remained locked in the cells and tissues of plants and animals. However, carbon is returned to the atmosphere as CO_2 through several pathways. First, and most direct, is cellular respiration, in which living plants and animals break down (oxidize) complex carbon compounds to release chemical potential energy for life processes and smaller carbon compounds, mostly CO_2:

$$C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$$

The CO_2 released is then available to plants for photosynthesis and the cycle continues.

Photosynthesis and respiration are generally closely balanced, reverse reactions. However, these processes have seen changes in steady state equilibria over millennia, resulting in vast "sinks" of oxygen (in the atmosphere) and carbon (in the lithosphere) from the cycle.

The second and most abundant release of CO_2 into the atmosphere is the aerobic decomposition of waste products and dead plants and animals by bacteria and certain fungi that obtain energy from the process. Note that oxygen is needed for this process, although anaerobic bacterial decomposition under suitable conditions can "recycle" carbon as methane, CH_4.

Finally, CO_2 is released into the atmosphere by combustion of organic matter, such as wood, grass, and paper (made from wood), which may have removed carbon from the cycle as organic compounds for a short time, as well as fossil fuels, whose carbon has been removed from the cycle a relatively longer time.

The carbon cycle is a two component cycle: between the atmosphere and lithosphere and between the atmosphere and hydrosphere, with the atmosphere functioning as the common reservoir. One study, using C-14, has estimated that a CO_2 molecule's residence time in the atmosphere is 10-15 years on average. It is important to note that vast amounts of carbon in the oceans are removed from biological processing as relatively unusable sediments, one of the major "carbon sinks." Although it has been estimated that 2.5×10^{12} tons of CO_2 cycle through the oceans and atmosphere about once every 300 years, carbon is locked up as insoluble carbonates, either as calcareous sediments or in the form of calcite, $CaCO_3$, or dolomite, $CaMg(CO_3)_2$, a much longer time within the ocean. A second long-range carbon sink consists of the fossil fuels, primarily found in the lithosphere.

Approximately one part in 10,000 of all organic substances have been removed from the carbon cycle over the ages, primarily as deep ocean sediments. These ocean sediments become buried in clay and sand and are thus removed from the oxygen needed for decay. Methane (CH_4, natural gas) is produced when the appropriate anaerobic bacteria are present. Otherwise the solid carbonaceous material is buried deeper, with proportionate increases in pressure and temperature, until disproportionation reactions begin, producing more methane and longer chain hydrocarbons (petroleum).

On land, woody plants from the Carboniferous Era—a warm, moist period occurring approximately 250-350 million years ago—underwent a somewhat similar metamorphosis. This woody material was composed of approximately 0-16% protein, 20-55% cellulose, 10-35% lignin, and approximately 10% inorganic salts by mass. Much of the protein was incorporated into body structures of aerobic and anaerobic microorganisms initially converting the material to peat, the first step in coal formation. Likewise, cellulose, a complex carbohydrate composed of long glucose strands, was converted to CO_2 and H_2O by similar metabolic processes. Most of what remained was lignin, a complex, three dimensional polymer more resistant to bacterial degradation than cellulose. Lignin accumulated in stagnant swamps under water, eventually compacting to form peat.

Further sedimentation and crustal upheavals gradually caused geologic metamorphism as peat was subjected to varying amounts of heat and pressure, driving off additional CO_2 and H_2O. As a result, deposits composed largely of carbon, with small quantities of hydrogen, oxygen, nitrogen, iron, and sulfur, were left behind as the substance we know as coal. The rank of a coal indicates how nearly it approaches both the composition and structure of graphite (Figure 7).

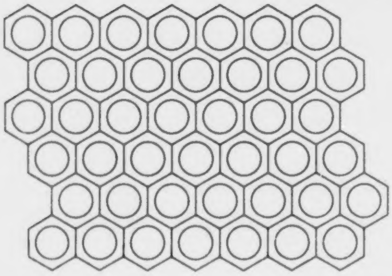


Figure 7. Structure of Graphite.

For example, under extreme conditions of temperature and pressure, anthracite coal forms when peat deposits become extensively aromatized, eventually condensing into polycyclic systems with very high carbon content, often as much as 95%, and with about 3% H and 2% O. The aromatic ring systems in anthracite may contain in excess of 30 condensed rings, developing structures that become more and more like graphite. The structures in Figures 7 and 8 illustrate the relationship between graphite and anthracite coal. Figure 9 shows that the structure of bituminous coal, a more open structure produced by less heat and pressure, contains fewer polyaromatic systems, indicating that bituminous coal is not as far along in the coalification process as anthracite coal.

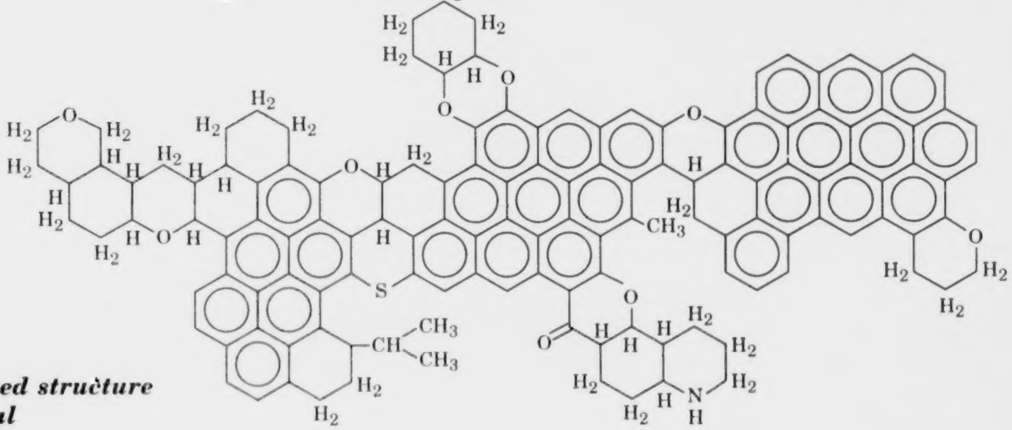


Figure 8. Proposed structure of Anthracite coal

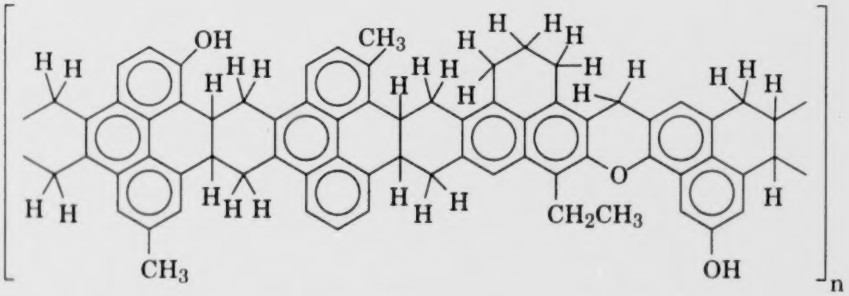


Figure 9. Proposed structure of bituminous coal.

		Analysis weight % before drying					
Fuel	Rank	State of origin	Moisture	Volatile matter	Carbon	Ash	Heat content (MJ/kg)
Anthracite	High	PA	4.4	4.8	81.8	9.0	30.5
Bituminous							
Low volatile		MD	2.3	19.6	65.8	12.3	30.7
Medium volatile		AL	3.1	23.4	63.6	9.9	31.4
High volatile		OH	5.9	43.8	46.5	3.8	30.6
Subbituminous		WA	13.9	34.2	41.0	10.9	24.0
		CO	25.8	31.1	34.8	4.7	19.9
Lignite (brown coal)		ND	36.8	27.8	30.2	5.2	16.2
Peat		MS	—	—	—	—	13.0
Wood	Low	—	—	—	—	—	10.4 - 14.1

Figure 10. Rank, composition, and fuel values of various United States coals.

Diamond, the hardest natural substance, is a product of this coal-forming process under the most extreme conditions. This tetrahedral, almost pure allotropic form of carbon, is produced from coal material at extremely high temperature and pressure deep within the earth. In fact, diamonds are usually found near areas of igneous intrusion near tectonic plate boundaries. Variation in color in diamonds is associated with mineral impurities.

There is much concern today about the increasing amounts of CO_2 released to the atmosphere by the use of fossil fuels and the "greenhouse effect" this use may be producing.

As shown in Figure 11, the major source of atmospheric CO_2 is not respiration by terrestrial plants and animals or the burning of fossil fuels, as might be expected, but fungal and microbial decomposition of dead organic matter. Extrapolating back from known data, scientists have found that CO_2 levels in the atmosphere seem to have maintained a fairly constant average at about 265 ppmv ± 30 ppmv (parts per million by volume) since life appeared on earth, with the exception of approximately the last 150 years. This relatively constant level over time seems to have been due to a CO_2/HCO_3^- equilibrium established between the atmosphere and the ocean surface. It should be noted, however, that there seems to have been some large natural fluctuations of CO_2 levels associated with climate changes in the past. Air samples trapped in northern hemispheric glaciers show that CO_2 levels dropped as low as 200 ppmv during the last Ice Age (15,000-30,000 yrs ago). This raises the question: was the CO_2 level low because of the cold climate (and hence the reduction of biotic respiration levels) or did a low CO_2 concentration level (cause unknown) lead to colder temperatures worldwide? There is disagreement among the experts, but most see some correlation between CO_2 concentrations and worldwide temperature.

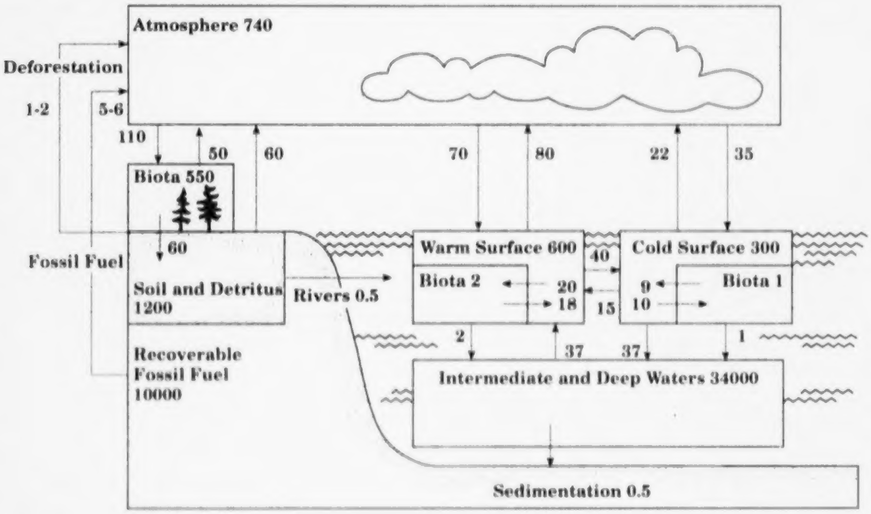


Figure 11. Fundamental features of the global carbon cycle. (Values shown X 10^{15} g carbon.)

As noted above, CO_2 levels since life began on earth have remained relatively constant until the last 150 years. This increase corresponds roughly to increased burning of fossil fuels during the Industrial Revolution. According to the U.S. EPA the concentration of CO_2 in the atmosphere has increased approximately 25% since the Industrial Revolution, and is currently increasing at an annual rate of 0.4%. Many hypothesize this increase should produce a gradual atmospheric warming called the "greenhouse effect," resulting in drastic climatic changes on earth.

In a true greenhouse, solar energy (visible light) is transmitted through the glass into the enclosure. Plants and other materials inside absorb this visible radiation and re-emit it in the form of lower energy infrared radiation (thermal energy). The glass, which transmits visible light, absorbs this infrared radiation and prevents its escape from the greenhouse (or automobile when the windows are up on a sunny day). As a result, the air becomes warm and is prevented from rising out of the greenhouse by the glass.

The atmospheric "greenhouse effect" is a partial misnomer, since there is no barrier in the atmosphere (like the glass) to prevent heat convection or wind cooling. However, several gases, the most important of which are H_2O and CO_2, absorb much of the long wavelength infrared radiation re-emitted from the earth's surface, causing the atmosphere to be warmer than it would be if all infrared radiation were re-radiated into space. Without the natural "greenhouse effect" the average temperature of earth would be around the freezing point of H_2O. As Figure 12 illustrates, an "atmospheric window" of infrared wavelengths around 10,000 nm are not absorbed, which allows some re-radiation into space and prevents the planet from overheating.

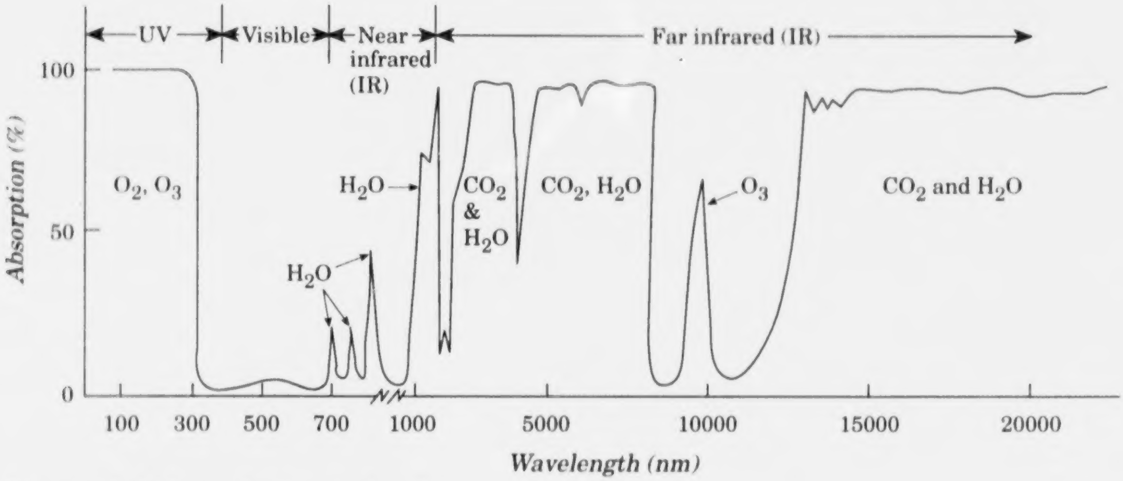


Figure 12. "Atmospheric window."

A good illustration of this effect (primarily caused by the "greenhouse" gas, H_2O) can be easily seen when comparing day/night temperature fluctuations in arid and humid areas of the planet. In arid locations, thermal energy (infrared radiation) emitted by the warmed surface, is rapidly lost to space since there is no water vapor to absorb it, and nighttime temperatures compared to daytime temperatures are very cold. In humid areas, water vapor absorbs some of the emitted thermal energy (infrared radiation), moderating the nighttime heat loss and producing less fluctuation in day/night temperatures (Figure 13).

It seems plausible to theorize that the steady increase of CO_2 in the atmosphere since the Industrial Revolution began to gradually cause greater absorption of re-emitted infrared radiation from the surface of the planet, with a resulting increase in global temperature. The effects of this increase could be disastrous. Many are aware that global warming, more rapid at the poles than equator, might result in a catastrophic rise in sea level as polar ice caps melt. However, few realize the possible effects on food production. Continuing increases in average worldwide temperature will shift wheat-growing zones towards the poles, away from areas of fertile to poorer land. Rates of desertification should increase in certain areas of the earth as a result of shifting rainfall patterns. Deep ocean upwelling of nutrients will be reduced as surface waters warm, producing widespread climatic changes that will decrease biological productivity.

Although increased concentrations of CO_2 should enhance photosynthetic rates in plants, there appears to be no evidence that there is increasing storage of carbon in the biota or detritus of the earth. Temperature data from two separate studies have shown that, during a sizeable part of the period since 1950, worldwide temperatures have actually decreased. In addition, there appears to be certain periods during each year when CO_2 levels are highest and lowest. As Figure 14 illustrates, peaks correspond to periods of low CO_2 demand (fall/winter) and troughs correspond to periods of high photosynthetic demand (spring/summer). It seems safe, therefore, to conclude that CO_2 levels do play a role in world climate, but that there are undoubtedly other factors that seem to be mitigating the effect, primary among them the interactions of the atmosphere-hydrosphere interface.

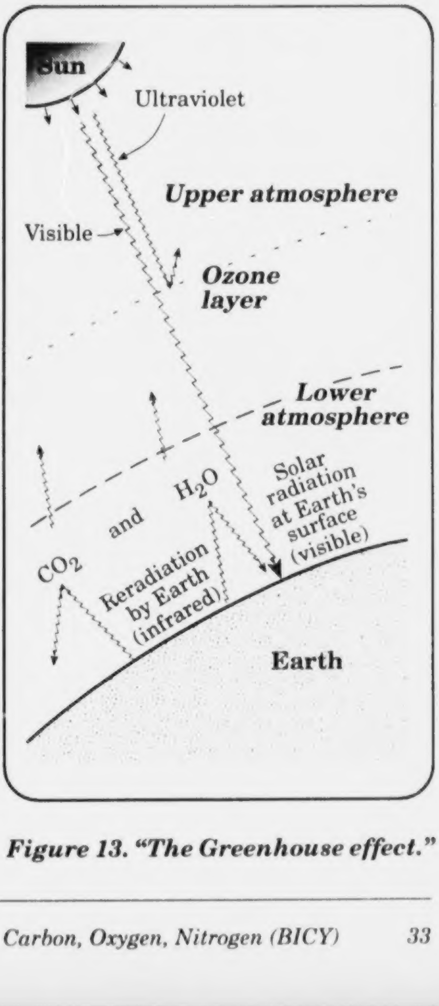


Figure 13. "The Greenhouse effect."

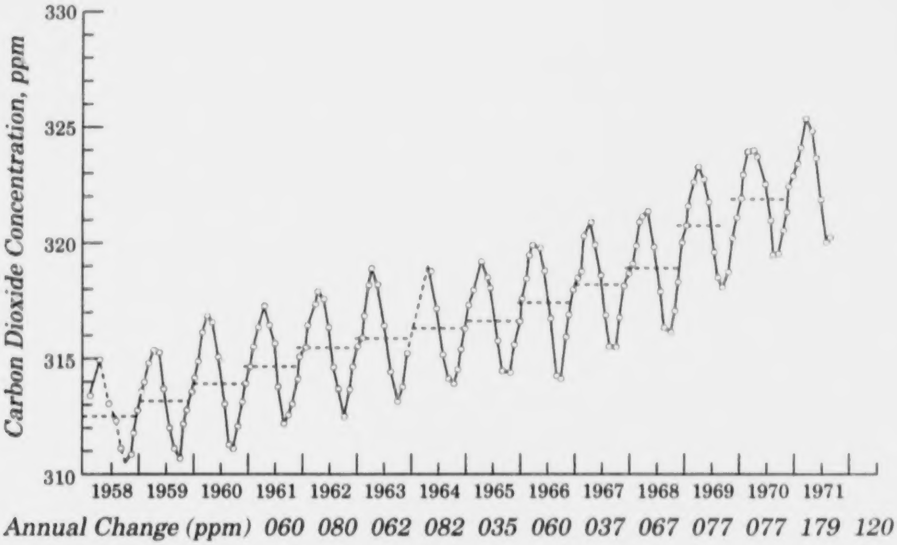


Figure 14. Buildup of carbon dioxide in the atmosphere.

Oxygen Cycle

The oxygen available to plants and animals has been produced over the past two billion years by photosynthesis in plants, primarily phytoplankton in the oceans. The oxygen produced by the earliest atmospheric organisms is locked away from the atmosphere in the banded iron formations deposited 2-3 x 10^9 years ago as Fe^{2+} was oxidized to Fe^{3+}. Oxygen was also used up in the oxidation of reduced atmospheric gases, like NH_3, CH_4, and H_2S. However, only with the evolution of protein molecules needed for oxygen transport did aerobic photosynthetic organisms spread through the hydrosphere, precipitating the remaining Fe^{3+} in seawater as iron(III) oxide, Fe_2O_3, and releasing oxygen as O_2 to the atmosphere. As an ozone layer developed, photoautotrophs (organisms capable of photosynthesis) began to range onto the land, producing larger and more varied populations, increasing the O_2 levels further.

Oxygen atoms (in CO_2) from the atmosphere become internalized as part of living organisms in carbohydrates or other organic substances. These are returned, along with additional oxygen atoms (O_2) from the atmosphere, as CO_2 during the respiration process. This cycle from photosynthesis to respiration and back to photosynthesis is continuous.

Oxygen is needed in respiration to break down carbohydrates in which carbon has been stored within the biota, either directly from photosynthesis (plants) or by acquisition from other organisms (animals). In this way oxygen is depleted from the atmosphere as carbon is released back to it. In certain cases where carbon was trapped away from oxygen (coal and petroleum formations), the oxygen released originally is not used in respiration, causing a surplus of O_2 to build up in the atmosphere. This build-up is believed to be the source of the present reservoir of approximately 10^{15} tons of oxygen in the atmosphere.

There is some concern today about depleting available oxygen by the burning of fossil fuels or the destruction of the tropical rain forests. However, J. Calvin Giddings (see *References*) proposes that only 1 part in 60 of the O_2 in the atmosphere would be consumed if all of the worldwide reserves of *recoverable* fossil fuels were burned. Likewise, he says that catastrophic destruction of the ocean's phytoplankton and land plants would produce "no serious dent in the O_2 supply, but it would terminate the ultimate source of all food and lead to rapid extinction of life."

It is important to note that the 10^{15} tons of oxygen in the atmosphere today represents only 1/21 of the oxygen evolved over geologic time. The rest has been used to oxidize H_2S, SO_2, NH_3, and CO emitted by volcanoes over the ages and solid iron and sulfur in the crust. Since this process continues today, the slow erosion of the global oxygen content must continue to be met by photosynthesis and organic sediment formation. This is one reason why students *should be concerned* about pollution of the oceans and rapid deforestation occurring worldwide. Though not immediately threatening, these processes cannot but impact the "oxygen reservoir" in time.

Nitrogen Cycle

In this cycle, free atmospheric nitrogen (N_2), is converted into soluble nitrogen compounds, then incorporated into plants and animals, and then returned to the atmosphere through organic decomposition.

Because the triple bond of molecular nitrogen, ($N{\equiv}N$), is so strong, most living things cannot use it directly. Atmospheric nitrogen must first be converted into a usable form by a process called fixation. Actually the term "fixation" applies to any gaseous substance that is condensed into a more usable form. (Thus atmospheric carbon dioxide is "fixed" by photosynthesis into the more usable form glucose.) The fixation of nitrogen can occur in three ways: biologically, atmospherically, and industrially.

Most natural nitrogen fixation occurs when nitrogen from the atmosphere is introduced into the soil where it is converted to ammonia compounds by soil bacteria such as *Rhizobium japonicum*. This process occurs on the roots of leguminous plants such as alfalfa, beans, and clover. (In Asia, nitrogen fixation occurs in rice paddies as a result of certain blue-green algae.)

The next step in nitrogen fixation is called "nitrification" and occurs when chemosynthetic bacteria oxidize ammonia compounds into nitrites (NO_2^-) and nitrates (NO_3^-), enriching the soil for plants, which use soluble nitrates to form plant proteins. When these plant proteins are eaten by animals, the nitrogen is removed as amino acids and used to build animal protein. When plant and animal matter decay, other bacteria break down amino acids from these organisms or their waste products and release ammonia compounds.

The nitrogen cycle is completed when anaerobic bacteria break down nitrates and other nitrogen compounds in the soil to release free nitrogen back into the atmosphere. Thus, the nitrogen cycle is dependent at each step upon microorganisms and could not proceed without them.

Nitrogen fixation occurs atmospherically when lightning and fires cause temperatures high enough to form oxides of nitrogen. These oxides are washed to earth by rain as nitrous (HNO_2) and nitric acids (HNO_3), adding the nitrate ion to both the soil and the waters of the earth.

Nitrogen fixation occurs industrially using the Haber process that combines nitrogen and hydrogen to form ammonia.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Ammonia is one of the top industrial chemicals produced in this country each year. A large amount of manufactured ammonia is converted to nitric acid in the Ostwald process and is then made into inorganic nitrates, such as NH_4NO_3, and used for fertilizer (see *Industrial Inorganic Chemistry*).

The various nitrogen cycles are illustrated in Figure 15.

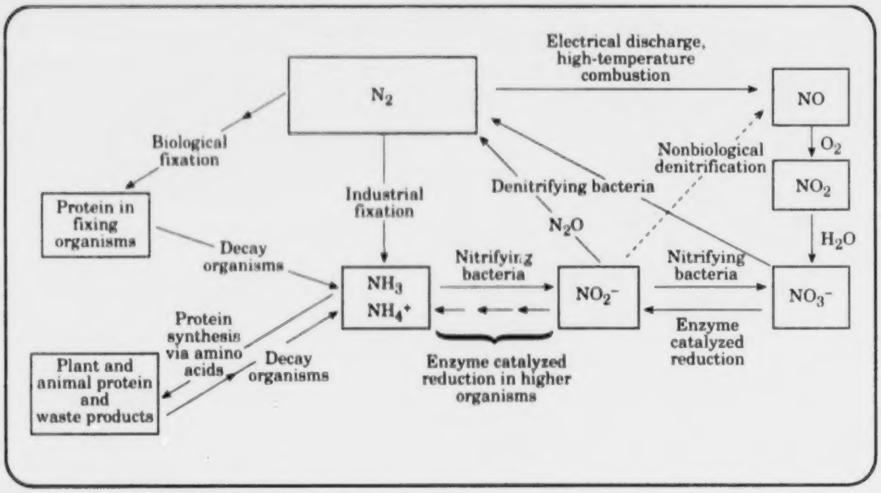

Electrical discharge, high-temperature combustion
N_2
NO
O_2
Biological fixation
Nonbiological denitrification
NO_2
Industrial fixation
Denitrifying bacteria
Protein in fixing organisms
N_2O
H_2O
Decay organisms
Nitrifying bacteria
Nitrifying bacteria
NH_3
NH_4^+
NO_2^-
NO_3^-
Protein synthesis via amino acids
Enzyme catalyzed reduction
Decay organisms
Plant and animal protein and waste products
Enzyme catalyzed reduction in higher organisms

Figure 15. Nitrogen cycles.

Nitrogen Oxides of Interest

Dinitrogen oxide (N_2O), a relatively unreactive oxide of nitrogen is produced by certain microbes. It was commonly used as an anesthetic known as "laughing gas." This compound reacts in the stratosphere to produce nitrogen gas and atomic oxygen.

$$N_2O(g) + h\nu \rightarrow N_2(g) + O(g)$$

Atomic oxygen can react with more N_2O to produce nitrogen and oxygen gas.

$$N_2O(g) + O(g) \rightarrow N_2(g) + O_2(g)$$

Nitrogen monoxide (NO), and nitrogen dioxide (NO_2) usually appear together in the atmosphere and are often expressed as NO_x. NO enters the atmosphere from natural sources, including biological processes and lightning. Large amounts enter the air from pollutants, especially as the products of fossil fuel combustion. Until recently, over half of NO_x in the atmosphere was produced by the automobile.

High temperatures, such as those associated with lightning, are required to form NO.

$$N_2(g) + O_2(g) \rightarrow 2NO(g)$$

Further reaction of NO with oxygen produces NO_2.

$$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$$

Nitrogen dioxide can undergo many reactions that involve oxygen atoms, oxygen gas, and ozone. NO_2 is ultimately removed from the air as organic nitrogen, nitrates, or nitric acid. These compounds form major sources of nitrogen for biological nitrogen cycles. The equation below shows how rain converts nitrogen dioxide to nitric acid, a component of acid rain.

$$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$$

NO and NO_2 differ in their properties and reactivity. NO is a colorless, odorless gas; NO_2 is a reddish-brown, toxic gas that causes severe respiratory problems in animals (including humans) and even death in high concentration. NO_2 is one of the major components of "photochemical smog." Photochemical smog occurs when ultraviolet light fuels the reaction of nitrogen dioxide with unburned hydrocarbons from exhaust of automobiles to produce ozone and other irritating chemicals causing health hazards in urban areas. NO_2 can also damage some metals and fade fabric dyes. (For additional information refer to the *Photochemistry* and *Industrial Inorganic Chemistry* modules.)

Since nitrogen oxide and nitrogen dioxide can also react with ozone, there is concern that adding more NO to the atmosphere might deplete the ozone layer.

$$NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$$

The ozone layer protects us from damaging ultraviolet radiation by absorbing certain wavelengths of light (240-300 nm) and preventing it from reaching earth. Supersonic planes, which travel in the stratosphere, emit considerable amounts of NO_x during flight. As the number of supersonic flights increase, there is concern that the amount of ozone in the stratosphere might be significantly diminished.

Language of Chemistry

atmosphere layer of air that surrounds the earth's surface. (The innermost layer of the atmosphere is the troposphere. It extends about 20 km [12 mi] above the earth and contains 95% of the air mass of the atmosphere.)

autotroph organism utilizing inorganic materials as food.

biogeochemical cycle mechanism by which elements, such as carbon, oxygen, and nitrogen are continuously reused by plants and animals.

biota all living organisms of the earth.

chlorophyll green pigment in plants that absorbs energy from the sun and catalyzes the reactions that produce carbohydrates.

deforestation excessive removal of trees from an area without adequate replanting.

detritus dead organic matter; excreted wastes and body remains of plants and animals.

fixation conversion of nitrogen into soluble nitrogen compounds, largely by certain bacteria in the soil but also by lightning.

nitrification oxidation of nitrites to nitrates by certain bacteria in the soil.

photochemical smog smog is a term originally coined from the words *smoke* and *fog*. It is currently used to describe a photochemical haze produced by the reaction of sun and atmosphere on certain compounds of nitrogen, sulfur, and carbon produced primarily by automobile exhaust.

photosynthesis series of reactions that takes place in plant chloroplasts to convert carbon dioxide and water in the presence of sunlight into carbohydrates and oxygen.

respiration series of cellular reactions in plants and animals that convert sugars and oxygen into carbon dioxide, water, and energy.

stratosphere part of the atmosphere extending from 20-50 km (12-30 mi) above the surface of the earth. It contains the ozone layer.

Common Student Misconceptions

1. **"The world's supply of oxygen is in danger of being used up."**

   In fact, if all of the recoverable fossil fuels were burned at one time, the oxygen supply would be reduced only by about 0.05%.

2. **"Nitrogen gas and oxygen gas readily react with each other."**

   The fact that these two gases stay separated in the atmosphere indicates that they are generally unreactive toward each other. In fact, nitrogen reacts with oxygen only at very high temperatures. Lightning is one of the few natural conditions that can generate enough energy for this reaction. The required temperature is also reached in boilers of fossil-fuel plants, automobile engines, jet engines, and large incinerators.

3. **"Photosynthesis requires a green plant."**

   Photosynthesis requires the presence of photosynthetic pigment, chlorophyll, found in green plants, but there are other pigments found in brown plants, red plants, and certain seaweeds that trap wavelengths of light that cannot be absorbed by chlorophyll alone.

4. **"Photosynthesis is a simple process."**

   Although we often show it as a series of 3 or 4 reactions, photosynthesis actually consists of 80-100 different, but interconnected, chemical reactions. Cellular respiration, essentially the reverse of photosynthesis, also consists of around 100 chemical reactions.

5. **"Carbon dioxide is the chief absorber of infrared radiation and is mainly responsible for the 'Greenhouse effect.'"**

   Water vapor absorbs infrared radiation over a broader range than CO_2 (see Figure 12), making it the atmospheric constituent most responsible for the "Greenhouse effect."

History: On the Human Side

An interesting historical note: **Professor Joseph Black**, a noted and very popular lecturer in Edinburgh in the 1770s, appears to be one of the first to have used this test for what he called "fixed air." In one of his lectures, he demonstrated that respiration produces the same "fixed air" that burning charcoal does:

> "Into this Glass syphon, I shall pour a quantity of Limewater... I now apply my mouth to the pipe, and suck in the common air through it; The fluid bubbles a little, but it is not altered in its transparency... But I now blow through it, and it becomes instantly muddy, the fixed air from the Lungs being attracted by the Lime, it loses its solubility and is precipitated."

Leonardo da Vinci (1452-1519) said, "Where flame cannot live, no animal that draws breath can live." *[A flame will be extinguished in* CO_2*.]*

Robert Boyle (1635-1691) conducted experiments to show that air has weight, and proposed that air contains a "vital quintessence" essential to animal life. *[This is oxygen!]*

Johann Becher (1635-1682) and **Georg Ernst Stahl** (1660-1734) proposed the "phlogiston theory" to account for burning. Things that burned readily contained more "phlogiston."

John Clayton (1657-1725) discovered "coal gas" by heating and distilling coal and wood.

Carl Wilhelm Scheele (1742-1786) and **Joseph Priestley** (1733-1804) independently prepared and described oxygen gas. Priestley discovered nitrous oxide (N_2O, dinitrogen oxide) in 1772.

Antoine Lavoisier (1743-1794) showed that a component of air was consumed when a substance burned in air. After observing the same effect with a mouse confined in air, he proposed that the reaction with air occurred in the lung of the animal as well. These ideas about burning brought about the downfall of the phlogiston theory.

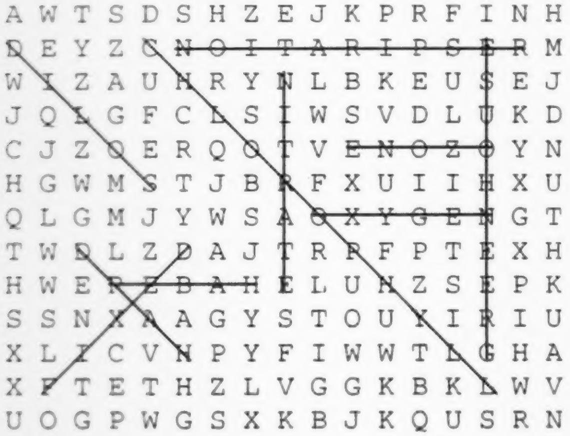
Daniel Rutherford (1749-1819) discovered nitrogen in 1772. He named the substance "noxious air" because it did not support life. The French word for nitrogen is "azote," meaning "without life."

Jan Ingenhousz (1730-1799) showed that plants, with the aid of sunlight, restored air to its "normal state," that is, provided oxygen, and proposed that plants receive their carbon dioxide from the air.

Nicolas Theodore de Saussure (1767-1845) explained the carbon cycle. He performed quantitative experiments in photosynthesis to show that plants extracted carbon dioxide from the atmosphere, that all the carbon in plants came from this gas, and that water was essential for photosynthesis.

HUMOR: ON THE FUN SIDE

1. Word Search (see *Appendix* for master copy)



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A W T S D S H Z E J K P R F I N H
D E Y Z C N O I T A R I P S E R M
W I Z A U H R Y N L B K E U S E J
J Q L G F C L S I W S V D L U K D
C J Z O E R Q O T V E N O Z O Y N
H G W M S T J B R F X U I I H X U
Q L G M J Y W S A O X Y G E N G T
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X L I C V N P Y F I W W T L G H A
X F T E T H Z L V G G K B K L W V
U O G P W G S X K B J K Q U S R N
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Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. A green plant's pigment.
2. Soil bacteria can oxidize ammonia to this substance.
3. This gas is a product of photosynthesis.
4. ______ effect, warming trend possibly due to trapping of infrared radiation by atmospheric gases.
5. The more usable form of a gaseous substance, *e.g.*, nitrogen.
6. This process combines N_2 and H_2 to form ammonia.
7. Adjective describing water that contains appreciable amounts of dissolved calcium and magnesium salts.

8. Cellular reaction involving conversion of sugar and oxygen into carbon dioxide and water.
9. The lithosphere is this part of our planet.
10. This allotrope of oxygen protects us from harmful ultraviolet radiation.

Answers: 1. CHLOROPHYLL 2. NITRATE 3. OXYGEN 4. GREENHOUSE 5. FIXED 6. HABER 7. HARD 8. RESPIRATION 9. SOLID 10. OZONE

2. See cartoons at end of module.

MEDIA

1. "The Fragile Planet" Videotape Series is published by Films for the Humanities and Sciences, P.O. Box 2053, Princeton, NJ 08543; (800) 257-5126.
   a. *Alterations in the Atmosphere*
   b. *The Heat Is On (Global Warming)*
   c. *The Ecology of the Forest*
   d. *Green Energy*
2. "Chemistry in Action" Videotape Series is published by Films for the Humanities and Sciences, P.O. Box 2053, Princeton, NJ 08543; (800) 257-5126.
   a. *Air Pollution Outdoor/Indoor*
   b. *Ways with Coal*
   c. *Out of the Air (1 and 2)*
3. The *World of Chemistry* videotapes "No. 17: The Precious Envelope," "No. 18: The Chemistry of the Earth," "No. 21: Carbon," and "No. 25: Chemistry and the Environment" are particularly appropriate for this module. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.
4. "Future Quest" series available from Hawkhill Associates, Inc., 125 E. Gilman, Madison, WI 53703; (608) 251-3934; (800) 422-4295; (608) 251-3924 (FAX).
5. Film/Filmstrip series available from Educational Images, Inc., Elmira, NY 14902; (607) 732-1090: "The Carbon Cycle," "The Nitrogen Cycle," "Carbon," and "Nitrogen."
6. Apple II computer diskettes available from Learning Arts, Wichita, KS 67201; (316) 636-9274: "Air Pollution," "Hothouse Planet" and "Photosynthesis."
7. Videodisc published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Earth's Atmosphere," "Depletion of Ozone," "The Hydrologic Cycle" and "The Carbon Dioxide Cycle," four chapters on *The World of Chemistry: Selected Demonstrations and Animations*: Disc II (double sided, 60 min.), Special Issue 4.
8. Written materials available from ICE – Institute for Chemical Education, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-3033 (voice) or (608) 262-0381 (FAX).
   a. Topics in Chemistry: *Acid Rain.*
   b. Topics in Chemistry: *Ozone–Chlorofluorocarbons and the Hole in the Ozone Layer.*
9. Pamphlets and brochures are available from the American Coal Foundation, 1130 17th St., N.W., Washington, DC 20036; (202) 466-8630.

Links/Connections

WITHIN CHEMISTRY

1. Elements and compounds
2. Properties of gases (solubility, *etc.*)
3. Acid-base reactions
4. Chemical reactions
5. Equilibrium
6. Oxidation-reduction
7. Organic and biochemistry

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Biology**: Photosynthesis, cellular respiration, microbiology, plant physiology
2. **Geology/geochemistry**: Agriculture, plant, and animal wastes
3. **Ecology/environmental science**: Pollution, global warming, deforestation, nitrogen oxide emission/photochemical smog control, combustion
4. **Business, economics, and politics**: Urban planning/development; cleanup of chemical dumping sites, oil spills; fossil fuel conservation; acid rain

TO THE CONTEMPORARY WORLD

Personal

1. **Careers**: Analytical chemist, biochemist, environmental chemist, field biologist, forester, geochemist, medicine, petroleum engineering, recycling engineer, research chemist, soil analyst, soil conservationist, teacher/professor, water chemist

Community

1. **Field trips**: Environmental testing laboratory, coal mine, petroleum refinery, limestone cavern, local weather station, botanical garden, local water works laboratory
2. **Knowledgeable individuals**: Agronomists, environmental scientists, recycling engineers, research chemists, water chemists, soil analysts, art conservators, airline pilots
3. Virtually every part of this module has a link to community issues, concerns, and problems. A discussion session with students highlighting these issues might be fruitful.

Other

1. The successful aquarist wishing to raise tropical fish must understand how nitrifying bacteria detoxify nitrogenous waste in a closed system. The biological "filter" afforded by the gravel bed changes NH_3 to NO_2^- and NO_2^- to NO_3^- through bacterial action. Also, circulation provided by a mechanical filter insures that CO_2 will evaporate from the surface of the water. (Idea contributed by Angie Matamoros.)

2. Secondary wastewater treatment relies on bacterial action - another chemical cycle. (Idea contributed by Angie Matamoros.)

3. The "biological pump" of ocean bacteria transports carbon from the surface to the deep ocean, thus enhancing the mixing of atmospheric carbon dioxide. The ocean chemistry of carbon dioxide is unusual and has links to thermodynamics, kinetics, equilibrium studies, irreversible processes, solubility of gases. See Sarmiento (1993) in *References*.

4. The "Biosphere 2" two-year closed system "experiment" in Arizona was an attempt to simulate a mini-set of biogeochemical cycles. What factors, both unintentional and eventually necessary, made this "experiment" a limited open system? See *Time* (September 30, 1991, pp. 65-66; September 24, 1990, pp. 72-73).

Extensions and Projects

1. Several very important scientists have contributed to the biogeochemical cycle, including the following:
    a. **Ernest Solvay**. His process for manufacturing washing soda ($Na_2CO_3 \cdot 10H_2O$) and baking soda ($NaHCO_3$) uses carbon dioxide (see *Alkali Metals* module).
    b. **Fritz Haber**. He devised a way to produce ammonia by directly combining nitrogen and hydrogen gases (see *Industrial Inorganic Chemistry* module).
    c. **Friedrich Wilhelm Ostwald**. He devised a three-step process to prepare nitric acid from ammonia (see *Industrial Inorganic Chemistry* module).
    d. **Melvin Calvin**. His research gave a clearer picture of photosynthesis. He discovered the "dark reaction" of photosynthesis, which is called the "Calvin Cycle."

    Research the lives and work of these famous scientists in your school library. Which ones received the Nobel Prize for their research? Add others to the list.
2. Carbon shares some properties with silicon. A possible "silicon-based" life form has often been considered, similar in many respects to the "carbon-based" life form as we know it. Have students speculate on the possibility of such a role for silicon. (Silicon chemistry is discussed in the modules *Chemistry of Rocks, Minerals and Gems* and *Materials Science.*)
3. In the 1950s, the famous Miller-Urey experiments proved that nitrogen-containing amino acids, the building blocks of protein molecules, could have been synthesized from electrical discharges in a primitive atmosphere. Have students do a library research project on the work and prepare a report. [See Miller, S. L., and Urey, H. C. (1959). Organic compound synthesis on the primitive earth. *Science*, *130*, 245.]
4. Nitrous oxide, N_2O, is sometimes called "laughing gas." It is used now as an anesthetic for some surgery and is a favorite among children's dentists. When it was first discovered, it was considered quite a novelty. Chemists carried bladders filled with this gas to parties, and it was made in school laboratories and passed around so that everyone could experience the pleasing and unpredictable effect of this gas. Have students do library research and prepare reports on the discovery and use of "laughing gas."

References

Module developed by Lee Summerlin, Patricia Owens, and Robert Davis, the Alabama team.

Alling, A., Nelson, M., Silverstone, S. (1993). *Life under glass: The inside story of Biosphere 2.* Oracle, AZ: The Biosphere Press.

American Heritage Dictionary of the English Language. (1980). *Cycle.* Boston, MA: Houghton Mifflin.

Bolin, B., and Cook, R. B. (Eds.). (1983). *Scope 21: The major biogeochemical cycles and their interactions.* New York, NY: Wiley.

A monograph—part of a detailed, technical series of articles edited and published on behalf of the International Council of Scientific Unions. Chapter 1 on "Interactions of Biogeochemical Cycles" is a valuable general source.

Borgford, C., and Summerlin, L. (1988). *Chemical activities: A sourcebook for the science teacher.* Washington, DC: American Chemical Society.

Broad, W. J. (1993, 5 October). Too rich a soil: Scientists find the flaw that undid the Biosphere. *The New York Times,* p. C1.

Calvin, M. (1969). *Chemical evolution.* New York, NY: Oxford University Press.

A discussion of molecular evolution towards the origin of living systems on the earth and elsewhere.

Cairo, R. (Ed.), and Dreyfus High School Chemistry Institute Team. (1988). *Chemistry of life.* Princeton: Woodrow Wilson National Fellowship Foundation.

"The Biochemical Elements" is an excellent summary of many elements and their natural cycles.

Chissus, D. R. (1989). *Greenhouse warming.* Presentation at the summer D.O.E. Teacher Research Associates Program project at Argonne National Laboratory, Argonne, IL.

Giddings, J. C. (1973). *Chemistry, man, and environmental change—an integrated approach.* San Francisco, CA: Canfield.

Chapter 6, "Atmosphere, Oxygen, and Global Air Pollution" introduces the concept of a carbon sink being responsible for the oxygen reservoir in today's atmosphere, as well as an excellent in-depth discussion of the greenhouse effect and climate.

Gymer, R. G. (1973). *Chemistry: An ecological approach.* New York, NY: Harper and Row.

Earlier work, updated and expanded in *Chemistry in the Natural World.*

Gymer, R. G. (1977). *Chemistry in the natural world.* Lexington, MA: Heath.

A general text intended for those interested in biogeochemical processes and the impact of various forms of pollution on them. See Chapter 6, "Biogeochemical Cycles" for a brief overview of the major cycles with many useful, thorough diagrams.

Likens, G. E. (Ed.). (1981). *Scope 17: Some perspectives of the major biogeochemical cycles.* New York, NY: Wiley.

A monograph—part of the detailed and technical series of reports published by the International Council of Scientific Unions.

Manahan, S. E. (1979). *Environmental chemistry*. Boston, MA: Willard Grant.

Very useful text written for undergraduates containing many simple and clear diagrams of elemental cycles and pathways in the environment. Chapter 11 on "Nature and Composition of the Atmosphere" is particularly well written.

Miller, G. T. (1988). *Living in the environment*. 5th Ed. Belmont, CA: Wadsworth.

Moore, J. W., and Moore, E. A. (1976). *Environmental chemistry*. New York, NY: Academic Press.

Excellent discussion in Chapter 3 of the chemical evolution of the earth's modern atmosphere and the appearance of life.

"Ozone." *Topics in Chemistry*, available from Institute for Chemical Education, U. of Wisconsin/Madison, Madison, WI.

A history of theories involving interactions of ozone and CFC's in the atmosphere.

Rambler, M. B., Margulis, L., and Fester, R. (Eds.). (1989). *Global ecology, towards a science of the biosphere*. Boston, MA: Academic Press.

Contains many simple diagrams, including one on the carbon cycle, and is one of the best sources for more recent data on the greenhouse gases.

Sarmiento, J. L. (1993, May 31). Ocean carbon cycle. *Chemical and Engineering News*, 71, 30-43.

Comprehensive article documenting the fate of anthropogenic carbon dioxide.

Schobert, H. H. (1989). The geochemistry of coal. I. The classification and origin of coal. II. The components of coal. *Journal of Chemical Education, 66*, 242-244, 290-293.

A two-part article on coal formation as a detour in the global carbon cycle. Written as a Chem I supplement.

Spiro, T. S., and Stigliani, W. M. (1980). *Environmental issues in chemical perspective*. Albany: State University of New York Press.

One of the best sources used—clear, descriptive presentation in Part 1 of chemical processes involved in coal formation. Part 2 contains an excellent discussion, with clear diagrams, of the role of ozone in ultraviolet shielding and as a greenhouse gas. Many chemical equations are presented to illustrate the depletion of O_3 naturally and by various pollutants.

Summerlin, L., and Ealy, J. (1985). *Chemical demonstrations: A sourcebook for the teacher*. Washington, DC: American Chemical Society.

Summerlin, L., and Borgford, C. (1987). *Chemical demonstrations: A sourcebook for the teacher* (Vol. 2). Washington, DC: American Chemical Society.

Towle, A. (1989). *Modern biology*. New York, NY: Holt, Rinehart, and Winston.

A general, introductory high school biology text.

Tzimopoulos, N. D., Metcalfe, H. C., William, J. E., and Castka, J. F. (1990). *Modern chemistry*. New York, NY: Holt, Rinehart, and Winston.

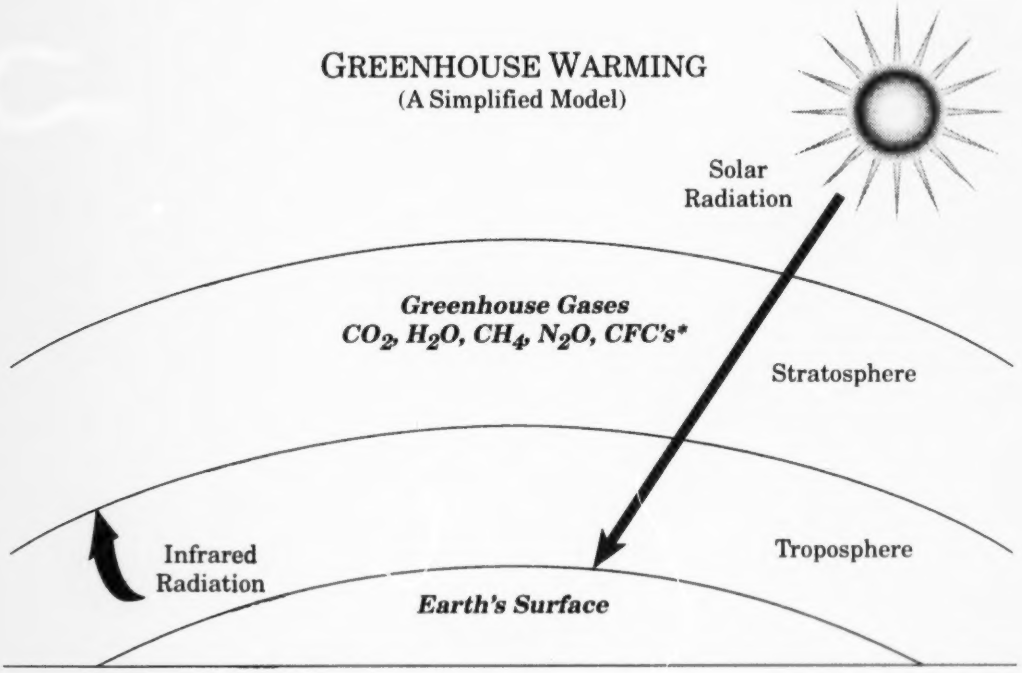
Probably the most widely used general high school chemistry text.

Appendix

- **Transparency Masters**
  1. Greenhouse Warming (A Simplified Model; Chissus, 1989)
  2. The Carbon Cycle (Gymer, 1977)
  3. Word Search
- **Humor**

GREENHOUSE WARMING

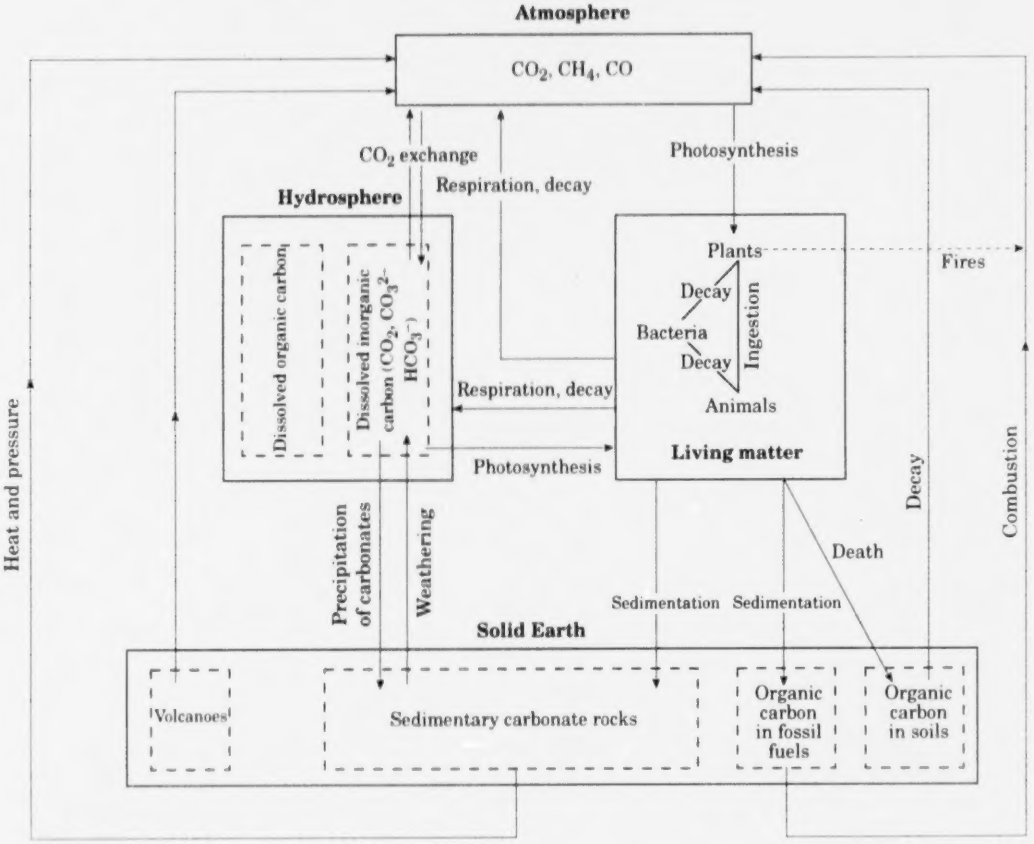
(A Simplified Model)


Solar
Radiation
Greenhouse Gases
CO_2, H_2O, CH_4, N_2O, CFC's*
Stratosphere
Infrared
Radiation
Troposphere
Earth's Surface

- Solar radiation (short wavelength) strikes the earth.
- Blanket of gases traps infrared energy (long wavelength) in the atmosphere.
- Without the greenhouse effect, the average temperature of the earth would be about freezing.

*Chlorofluorocarbons

The Carbon Cycle

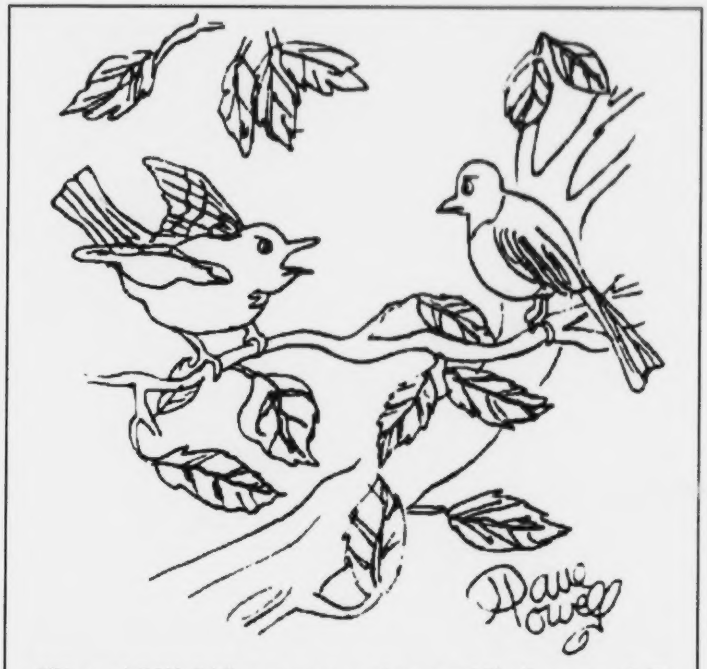

Atmosphere
CO_2, CH_4, CO
CO_2 exchange
Respiration, decay
Photosynthesis
Hydrosphere
Dissolved organic carbon
Dissolved inorganic carbon (CO_2, CO_3^{2-}, HCO_3^-)
Plants
Fires
Decay
Bacteria
Decay
Ingestion
Animals
Respiration, decay
Photosynthesis
Living matter
Heat and pressure
Precipitation of carbonates
Weathering
Sedimentation
Sedimentation
Death
Decay
Combustion
Solid Earth
Volcanoes
Sedimentary carbonate rocks
Organic carbon in fossil fuels
Organic carbon in soils

Word Search

A	W	T	S	D	S	H	Z	E	J	K	P	R	F	I	N	H
D	E	Y	Z	C	N	O	I	T	A	R	I	P	S	E	R	M
W	I	Z	A	U	H	R	Y	N	L	B	K	E	U	S	E	J
J	Q	L	G	F	C	L	S	I	W	S	V	D	L	U	K	D
C	J	Z	O	E	R	Q	O	T	V	E	N	O	Z	O	Y	N
H	G	W	M	S	T	J	B	R	F	X	U	I	I	H	X	U
Q	L	G	M	J	Y	W	S	A	O	X	Y	G	E	N	G	T
T	W	D	L	Z	D	A	J	T	R	P	F	P	T	E	X	H
H	W	E	R	E	B	A	H	E	L	U	H	Z	S	E	P	K
S	S	N	X	A	A	G	Y	S	T	O	U	Y	I	R	I	U
X	L	I	C	V	H	P	Y	F	I	W	W	T	L	G	H	A
X	F	T	E	T	H	Z	L	V	G	G	K	B	K	L	W	V
U	O	G	P	W	G	S	X	K	B	J	K	Q	U	S	R	N

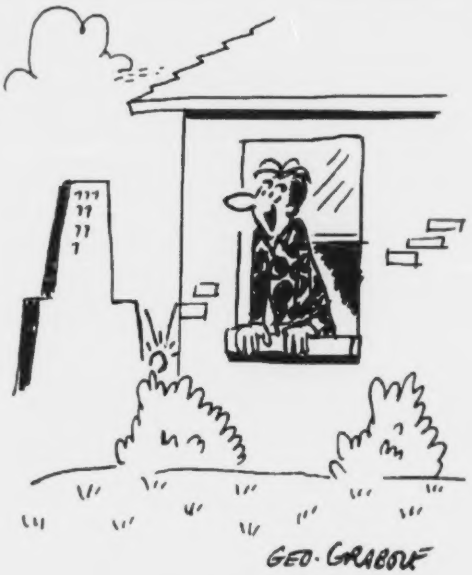
Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. A green plant's pigment.
2. Soil bacteria can oxidize ammonia to this substance.
3. This gas is a product of photosynthesis.
4. ______ effect, warming trend possibly due to trapping of infrared radiation by atmospheric gases.
5. The more usable form of a gaseous substance, *e.g.*, nitrogen.
6. This process combines N_2 and H_2 to form ammonia.
7. Adjective describing water that contains appreciable amounts of dissolved calcium and magnesium salts.
8. Cellular reaction involving conversion of sugar and oxygen into carbon dioxide and water.
9. The lithosphere is this part of our planet.
10. This allotrope of oxygen protects us from harmful ultraviolet radiation.



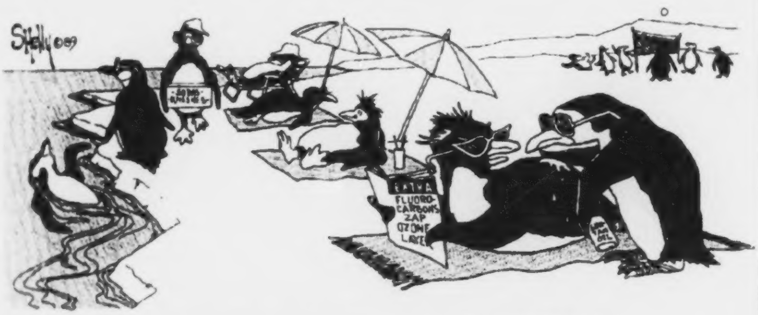
"Oh yeah? Well I live a more endangered life than *you* do!"

CHEMTECH, May 1980, p. 267. © American Chemical Society.
Reprinted with permission.


GEO. GRABOLF

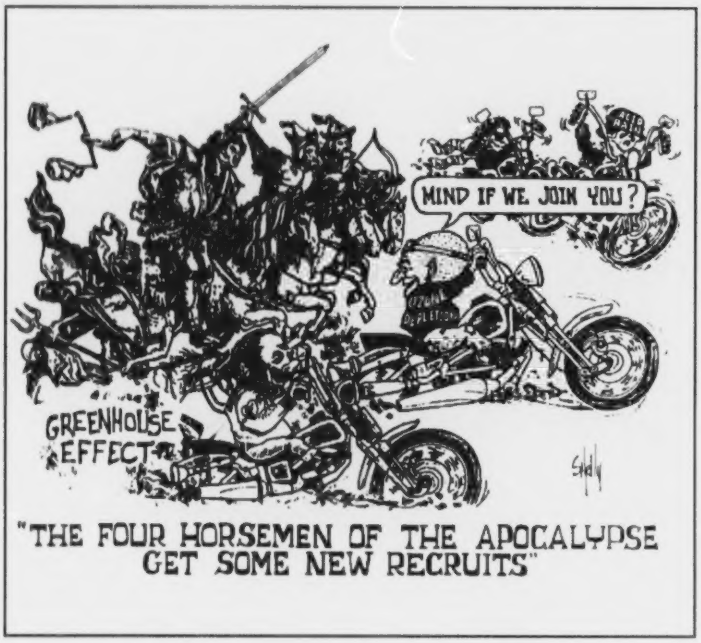
"Good morning, all of you out there in the ecosystem!"

CHEMTECH, November 1973, p. 686. ©American Chemical Society. Reprinted with permission.

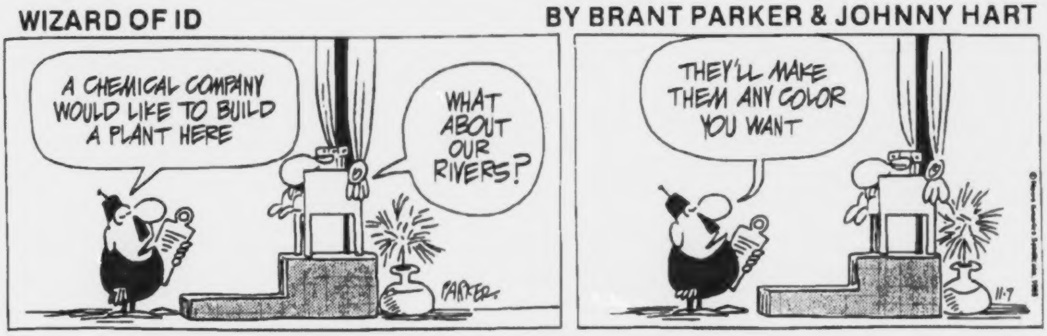

SHELLY ©89
EXTRA
FLUORO-CARBONS ZAP OZONE LAYER

OZONE PROBLEM? WHAT OZONE PROBLEM?

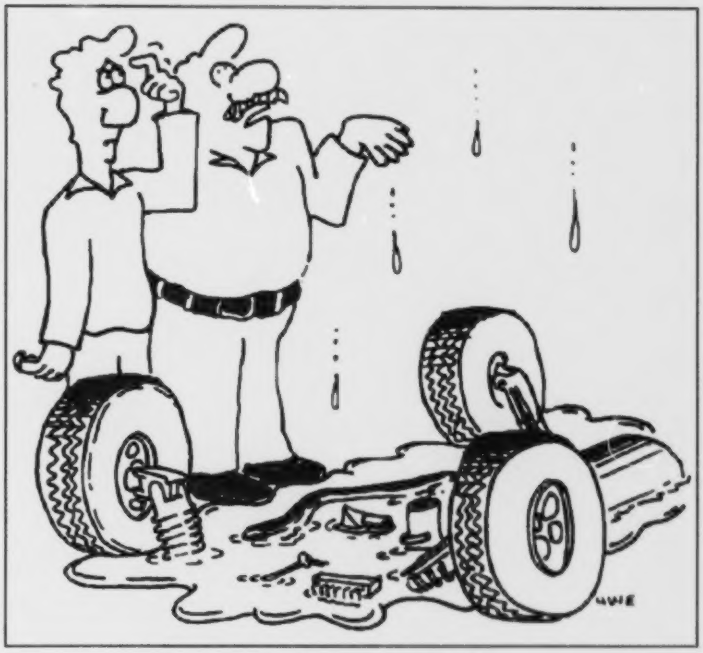
Cartoon by Shelly Fischman. Reprinted with permission.


MIND IF WE JOIN YOU?
OZONE DEPLETION
GREENHOUSE EFFECT
Shelly
"THE FOUR HORSEMEN OF THE APOCALYPSE GET SOME NEW RECRUITS"

Cartoon by Shelly Fischman. Reprinted with permission.

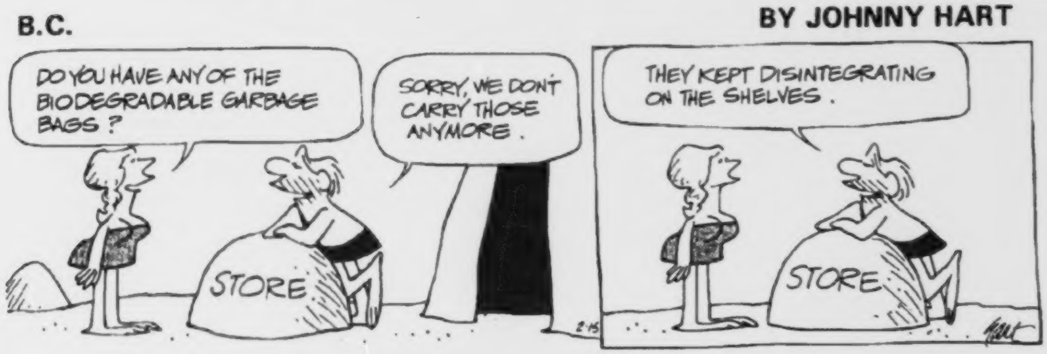

WIZARD OF ID
BY BRANT PARKER & JOHNNY HART
A CHEMICAL COMPANY WOULD LIKE TO BUILD A PLANT HERE
WHAT ABOUT OUR RIVERS?
PARKER
THEY'LL MAKE THEM ANY COLOR YOU WANT
11-7

By permission of Johnny Hart and Creators Syndicate, Inc.



MUST BE ACID RAIN !

CHEM 13 NEWS, January 1980, p. 1. Reprinted with permission.


B.C.
BY JOHNNY HART
DO YOU HAVE ANY OF THE BIODEGRADABLE GARBAGE BAGS ?
SORRY, WE DON'T CARRY THOSE ANYMORE .
THEY KEPT DISINTEGRATING ON THE SHELVES .
STORE
STORE
2-15
hart

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A SourceBook Module

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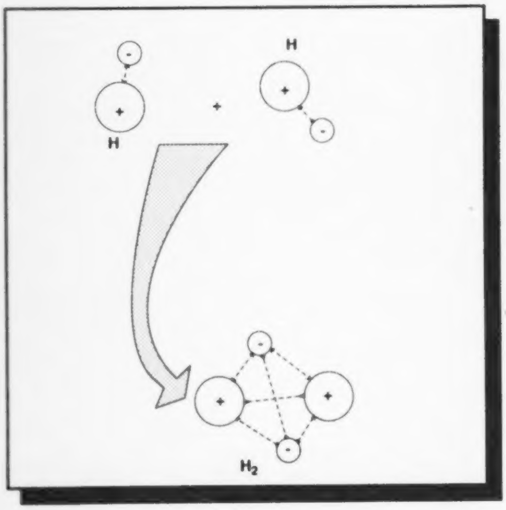
CHEMICAL BONDING



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The properties of elements can generally be explained in terms of electron occupancy at the highest energy level in the atom. These electrons are usually called *valence electrons*. Elements whose atoms have relatively few valence electrons are typically metals, while those with relatively large numbers of valence electrons are typically nonmetals. This is particularly true for the representative elements (main group elements). For example, alkali metal and alkaline earth elements are typically metals whereas oxygen family elements and halogens are typically nonmetals.

Atoms form molecules and gain stability (lowered potential energy) through covalent bonding by sharing one or more electron pairs. Electron pair sharing is typical of nonmetals and is best illustrated for the hydrogen molecule, a case of "equal sharing." Two isolated hydrogen atoms have only one electron each. As the atoms approach each other, the electron of one atom is attracted by the nucleus of the other atom, and *vice versa*. This mutual attraction by two nuclei for an electron pair gives rise to the covalent bond, with an excess of attractive forces overcoming repulsive forces. Only "like" atoms (*e.g.*, same electronegativity) form such equal-sharing, nonpolar bonds. Most covalent bonds are not "equal sharing" since there is frequently a difference in electronegativity. In these cases the bond is called a *polar covalent* bond. This means that the bond has a positively charged and negatively charged end. However, the charges are not nearly as large as the charges on ions in ionic solids. It is common practice to show the bonding in molecules in terms of Lewis-dot formulas, named in honor of G. N. Lewis. In these formulas, the elemental symbols represent the nucleus and all the electrons except valence electrons. The valence electrons are represented as dots—one dot means one electron, two dots two electrons, and so on. The structures are consistent in most simple cases with the octet rule.

Now consider the opposite extreme. When a metal and nonmetal form a compound, often the nonmetal attracts electrons more strongly than the metal (*i.e.*, it has a larger electronegativity). In such a case the electron pair is "taken over" by the more electronegative atom to form a negatively charged ion. The metal atom, by virtue of losing an electron, acquires a positive charge to form a positively charged ion. This is essentially what happens between an alkaline earth metal and a halogen. If two electrons are "transferred" as between an alkaline earth element and an oxygen family element, then 2+ and 2– ions are formed. In such cases the bonding is called *ionic* bonding and the stability (lowered potential energy) is due to the mutual attraction between oppositely charged ions in the solid crystals these compounds form.

The bonding categories considered thus far—covalent and ionic—explain the structure and properties of many substances. But metals, covalent network solids, and molecular solids are somewhat unique and require additional attention.

Bonding in metals, called *metallic* bonding, involves valence electrons. These electrons are loosely held by any one atom and collectively form a "sea of valence electrons" that can be used to explain many metallic properties, *e.g.*, metallic luster, malleability, electrical conductivity, *etc.* The electrons are loosely held since each atom has several unoccupied valence orbitals; it is relatively easy for the electrons to move about. In this manner the electrons allow atoms to slide past each other and be "worked" (hammered) into shapes and drawn into wires (evidence of malleability and ductility). The mobile electrons in appropriate circumstances move and conduct electricity and heat. The metallic bond is the third and last of the generally recognized types of chemical bonds (see *Materials Science* module).

All atoms in *covalent network* solids are covalently bonded. The high melting points, hardness, and other properties are explained by covalently bonded atoms forming three, two, and one dimensional networks, *e.g.*, quartz, mica, and asbestos (see *Materials Science* module).

Many properties of molecular solids and liquids are due to *van der Waals forces*. These forces are relatively weak forces between neighboring molecules either due to momentary dipoles or due to electrostatic interactions among polar molecules. Van der Waals forces are commonly classified as either *London forces* (between molecules of hydrogen and carbon dioxide for example) or *dipole-dipole* interactions (between molecules of hydrogen chloride and chloroform for example). *Hydrogen bonding*, a special case of dipole-dipole interaction, occurs between hydrogen atoms bonded to nitrogen, oxygen, or fluorine atoms in one molecule and a nonbonding electron pair on the nitrogen, oxygen, or fluorine atom in an adjacent molecule (as in water and ammonia, for example). In some cases hydrogen bonds can form *within* the same molecule, depending on its geometry and composition.

PLACE IN THE CURRICULUM

Bonding logically follows development of a detailed atomic model of matter including electron configurations and a consideration of chemical periodicity. The topic of bonding may be treated as a whole or may be partitioned into several segments. Personal preference or the organization of the adopted textbook may dictate order of coverage. In any case, ionic and covalent bonding should be treated, while covalent networks, van der Waals forces, and metallic bonding may be deferred until condensed phases are taught.

CENTRAL CONCEPTS

1. The electrons in the highest occupied energy level in a ground-state atom determine the atom's chemical properties and are called valence electrons.
2. Atoms may form molecules or aggregates and gain stability by sharing an electron pair, giving rise to an electrostatic attraction, thus forming a covalent bond.
3. Metals and nonmetals may form a stable structure by transferring electrons from the metal atom to the nonmetal atom. This transfer forms oppositely charged particles called ions; the resulting structure is held together by ionic bonds.
4. Valence electrons are arranged around the kernels (the nucleus and electrons other than the valence electrons) of atoms, forming molecules. Valence electrons can be represented as dots, yielding structures (Lewis-dot structures) consistent with the octet rule.
5. Carbon, nitrogen, oxygen, and sulfur atoms frequently share two or three electron pairs, forming double or triple covalent bonds consistent with the octet rule.
6. Like atoms (*i.e.*, same electronegativity) form "equal-sharing" covalent bonds, called nonpolar bonds. Atoms with different electronegativities do not equally share electron pairs; they form polar covalent bonds.
7. When two isolated atoms capable of bonding approach each other, the potential energy of the system decreases due to attraction and is a minimum at some equilibrium distance (called the bond distance) separating the atoms. If the atoms attempt to approach more closely, the attraction changes to repulsion rapidly; hence the potential energy of the system also rises rapidly.

8. Positively charged atomic kernels in metals are held in closely packed crystal structures by mobile valence electrons.
9. Covalent network solids involve atoms covalently bonded in three, two, or one dimensional networks. The primary covalent bonds give these substances many of their properties.
10. Intermolecular (as opposed to intramolecular) attractions are responsible for aggregates of molecules that become solids or liquids. The weakest attractions are collectively called van der Waals forces. London forces, the weakest of these forces, are caused by momentary fluctuations of electron distribution symmetry in the atoms. The strength of theses forces depends on polarizability of the electron distributions and surface area of the molecules. Although London forces are the weakest of the van der Waals forces, they may be considered the most important, because they occur between every atom or molecule regardless of whether or not any other forces are present. Dipole-dipole interactions occur when polar molecules are attracted to each other. The strongest dipole-dipole interactions occur when a hydrogen atom, covalently bonded to nitrogen, oxygen, or fluorine, is attracted to an unshared electron pair of a similar highly electronegative atom in the same or adjacent molecule. This is called hydrogen bonding (see *Condensed States* module).

RELATED CONCEPTS

1. The Rutherford (nuclear) model of the atom and the quantum mechanical model of the atom in terms of electron configurations form the conceptual basis for chemical bonding.
2. There are four known kinds of forces: gravitational, electromagnetic, weak nuclear, and strong nuclear forces. Of these, bonding is determined only by electromagnetic forces.
3. Gravitational and electromagnetic forces obey an inverse square law. Coulomb's law, of interest in bonding, illustrates this: $F = \frac{q_1 q_2}{d^2}$

   In this expression, F is the coulombic force, q is the magnitude and sign of the two charges, d is the distance separating them, and k is the proportionality constant.
4. There is a general tendency for all matter to change to the lowest potential energy possible during chemical changes. A general knowledge of this tendency, energy units, and how energy changes are measured are helpful in understanding chemical bonding.
5. Concepts of atomic structure should be developed before considering bonding.
6. Concepts of periodicity should be developed prior to considering this topic.

RELATED SKILLS

1. Write electron configurations for common elements.
2. Decide, on the basis of electronegativity differences, bond polarity and bond type (*i.e.*, ionic or covalent).
3. Use and draw vectors to represent polar covalent bonds.

PERFORMANCE OBJECTIVES

After completing their study of chemical bonding, students should be able to:

1. determine the number of valence electrons in an atom from its position in the Periodic Table.
2. draw Lewis-dot structures of the representative elements.
3. state and apply the octet (duet) rule.
4. describe the role of noble gas electron configurations in ion formation.
5. describe the formation of cations and anions using orbital diagrams, Lewis-dot formulas, and standard ion formulas.
6. state the characteristics of an ionic bond and recognize compounds having ionic bonds.
7. relate the properties of ionic substances to ionic bonding.
8. explain the electrical conductivity of melted and of aqueous solutions of ionic compounds.
9. describe the metallic bond in terms of vacant valence orbitals and loosely bound valence electrons. Relate some properties of metals to the type of bonding present.
10. describe the formation of a covalent bond between two nonmetallic elements.
11. draw Lewis-dot formulas for simple covalent molecules containing single, double, and triple covalent bonds.
12. define polar covalence and use electronegativity values to determine whether a bond is nonpolar covalent, polar covalent, or ionic.
13. describe van der Waals forces (London forces and dipole-dipole forces) and hydrogen bonds.
14. explain why molecules are more stable than separated atoms.
15. predict how many bonds may be formed by atoms of representative elements.
16. predict simple physical properties (*e.g.*, melting point) based on the kind of van der Waals forces present.
17. predict simple properties of covalent network solids on the basis of covalent bonding in elements, *e.g.*, diamond and graphite, and in other examples of network solids, *e.g.*, silicates.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Physical Properties and Chemical Bonding in Solids

Introduction

In this activity you will investigate selected properties of four types of solids—ionic solids, molecular solids, covalent network solids, and metals. The properties studied will allow you to distinguish among the four types of solids. These properties will also allow you to determine the type of bonding in an unknown solid.

Recall that ionic solids have ions located at regular, repeating crystal lattice sites with bonding among ions being primarily electrostatic. This type of bonding is essentially nondirectional; ions arrange themselves in sites to form a crystal structure that maximizes attractions and minimizes repulsions. The electrostatic forces are strong, giving rise to large lattice energies—it takes a large quantity of energy to disrupt the crystal.

By contrast, covalent molecular solids have molecules held in place in a crystal lattice either by weak dispersion forces or dipole-dipole forces. These forces allow the crystal to be disrupted with a much smaller energy input than is the case for ionic solids.

Covalent network solids contain only covalent bonds, but the bonding is quite different between adjacent atoms in the crystal. These bonds consist of primary covalent bonds—strong chemical bonds. This kind of bonding in one, two or three dimensions gives rise to strong structures as in quartz, mica, and asbestos. It takes a very large energy input to disrupt the crystal lattice, leading to very high melting points.

Bonding in metals is quite different from that of other classes. Metallic bonding is electrostatic as in ionic solids, but the attraction is between valence electrons and the positively charged metal atom kernels. This creates nondirectional bonding; electrons are not strongly associated with any one particular atomic kernel. Thus the electrons are rather mobile, accounting for the properties of metals.

Since substances used in the activity are fairly typical of their classes, it is possible to generalize and correlate the structures and properties of the substances with the bonding involved.

Purpose

To investigate some physical properties of solids containing ionic bonds, covalent bonds, van der Waals forces, and metallic bonds to learn how to discriminate among them on the basis of properties correlated with their bonding forces.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Lauric acid is generally regarded as safe; however, as with all substances, exercise care while handling it.
3. Cyclohexane is highly flammable and is toxic by ingestion, inhalation, and absorption through the skin. Cyclohexane should be kept at least three meters from any open flame.

4. The electrical conductivity apparatus is a potential source of electric shock.
5. Volatile substances should be used in a fume hood.
6. Place the used cyclohexane in the container provided. Place the used lauric acid and other solids in the designated container(s).

Procedure

1. Place 1-g samples of sodium chloride, NaCl; lauric acid, $CH_3(CH_2)_{10}COOH$; silicon dioxide, SiO_2 (sand); and iron filings in separate evaporating dishes. Test hardness by rubbing each solid with the bottom of a plastic pen or stirring rod.
2. Cautiously smell each substance by fanning any vapors with your hand to your nose. An odor implies some volatility.
3. Place pea-sized samples of NaCl, lauric acid, silicon dioxide, and iron filings in separate small test-tubes. Test the melting point of each in a boiling water bath. Then place any of the substances that *did not melt* in a clean, small evaporating dish. In turn, place each evaporating dish on a wire gauze or wire triangle and heat directly with a burner flame. As soon as the solid melts, remove the flame. Do not heat any substance longer than five min. Remember that the heated substances require time to cool before disposal. Handle hot evaporating dishes with tongs and place on a ceramic pad to cool.
4. Arrange a test-tube rack with four pairs of test-tubes. Add 0.5 g NaCl to each test-tube in the first pair. Do likewise with the other solids in the remaining three pairs of test-tubes. Add 5-mL distilled water to the first of each pair of test-tubes. Add 5-mL cyclohexane to the second of each pair of test-tubes. Stopper and shake the test-tubes noting any evidence of dissolving of the solid.

NOTE: Steps 5 and 6, at the option of your teacher, can be done as a demonstration.

5. Test the conductivity of solid NaCl, lauric acid, SiO_2, and iron by placing a sample in an evaporating dish and touching the solid to the wires of a low-voltage conductivity tester.
6. Test the conductivity of distilled water and cyclohexane, distilled water solutions, cyclohexane solutions, and any solvent/solid mixtures of the four substances. Place the solutions and/or mixtures in a 50-mL beaker and immerse the electrodes. Rinse the electrodes with the solvent before and after each test.
7. Obtain an unknown solid and attempt to identify its bond type by repeating the tests suggested above.
8. Dispose of materials as earlier directed in the "general safety considerations" and/or as directed by your teacher. Thoroughly wash your hands and clean all equipment prior to leaving the laboratory.

Data Analysis and Concept Development

The properties of the four classes of substances studied showed marked differences. These differences are explained in terms of the bonding that holds the substances together in the solid state. Study the data table and list typical properties expected for the four kinds of substances, *e.g.*, order of melting, hardness, solubility. Remember that water is a polar solvent while cyclohexane is a nonpolar solvent. Identify the unknown as being either ionic, covalent molecular, covalent network solid, or metal. Then answer the *Implications and Applications* questions that follow.

Data Analysis Summary

Ionic solids:

Covalent molecular solids:

Covalent network solids:

Metals:

Classification of unknown:

Implications and Applications

1. Explain the odor data in terms of bond types. Do you think that a substance can be relatively volatile and not have an odor? Give an example. How could the sensitivity of one's nose affect the observations made?
2. Explain differences in melting points in terms of types of bonding in the four substances.
3. Does a comparison of melting points give any information about the relative strengths of ionic and covalent bonds? Explain.
4. Are intermolecular or intramolecular forces greater for lauric acid? Explain your answer.
5. How can the results of the solubility tests be explained? What is the major difference in the nature of water and cyclohexane as solvents?
6. Explain the classification of the unknown as ionic solid, covalent molecular solid, covalent network solid, or metal.

Activity 1 was adapted from Experiment B22 in the following source:

Brooks, D. W. (Producer). (1989). *Doing chemistry* [Videodiscs, computer program, and supporting written materials]. Washington, DC: American Chemical Society.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Physical Properties and Chemical Bonding in Solids

Major Chemical Concept

To a large extent, the physical properties of a solid substance are determined by the type of bonding that holds molecules, atoms, or ions together. Solids may be classified as ionic, molecular, metallic, or covalent network solids. The type of bonding affects physical properties of a solid and may be determined by studying its physical properties.

Level

This activity is suitable for basic, general, and honors students. Only the part involving sodium chloride and lauric acid is recommended for basic students, although doing the remainder as a guided teacher demonstration might be useful for these students.

Expected Student Background

Students should have some background in ionic and covalent bonding. Previous instruction in van der Waals forces, bonding in covalent network solids, and metallic bonding is not important, since this laboratory activity provides a natural lead into a discussion of these types of bonding.

Time

Teacher preparation requires about 40 min. Students should be able to complete the activity in about 50 min.

Safety

Usual safety precautions include safety goggles, gloves, aprons, and availability of a fume hood. Students must be monitored during the conduct of the activity, especially if they do the conductivity testing. It is recommended that you perform the conductivity testing for students if a device powered by 120 V is used. If low-voltage battery-powered testers are available, then students can complete this. Monitor students to insure that no open flames are present while cyclohexane is being used. Have a suitable fire extinguisher available.

Materials (For 24 students working in pairs)

Nonconsumables

60 Evaporating dishes
96 Test-tubes, 18- x 150-mm, rubber or cork stoppers to fit
12 Test-tube racks
12 Ringstands, iron rings, triangles, wire gauzes, burners
12 Crucible tongs
12 Low-voltage conductivity testers
12 Beakers, 50-mL (100-mL beakers may be used)
12 Graduated cylinders, 10-mL
12 Spatulas
12 Plastic pens or other devices for crushing solids
12 Beakers, 250-mL
Safety goggles (one per student)
Fume hood

Consumables

Sodium chloride, $NaCl$, 48 g
Silicon dioxide, SiO_2 (silica gel or pure sand), 48 g
Lauric acid, 48 g
Iron filings, 48 g
Cyclohexane (or charcoal lighter fluid), 360 mL
Distilled water

Advance Preparation

1. It is best to have the solids and liquids available for dispensing in numerous small containers convenient to students. Dispense cyclohexane in a fume hood.
2. If you have not performed this activity before, plan to do so prior to having students do it.
3. Make waste containers available and be prepared to give specific disposal directions to students prior to the activity.
4. Check the working order of the fire extinguisher, fire blanket, safety shower, and eye wash.
5. Safety goggles must be available.

Pre-Laboratory Discussion

Review safe laboratory procedures and general safety rules. Demonstrate the safe handling of supplies and equipment used. Give students specific directions about disposal of lauric acid and cyclohexane. Tell students to stop heating a substance once it has melted. Also tell students that if they experience difficulty in melting NaCl in an evaporating dish, to try placing a small amount in a test-tube and heating the test-tube directly over the hottest part of the burner flame. Caution students not to use flames near cyclohexane.

Help students organize a suitable data table, possibly as shown here.

Sample Data Table

Observations

Property	**NaCl** (Ionic)	**Lauric acid** (Molecular)	**SiO_2** (Covalent)	**Fe** (Metallic)
Hardness				
Volatility/odor				
Melting				
Water solubility				
Cyclohexane solubility				
Conductivity/solid				
Conductivity/water solution				
Conductivity/ cyclohexane solution				

Anticipated Student Results

The following table presents sample data for this activity. Student groups will produce some variations, but data should resemble the sample data.

Sample Data Table

Observations

Property	NaCl (Ionic)	Lauric acid (Molecular)	SiO_2 (Covalent)	Fe (Metallic)
Hardness	Brittle	Soft	Hard	Hard
Volatility/odor	None	Fatty	None	None
Melting	Barely	Easily	No	No
Water solubility	Soluble	No	No	No
Cyclohexane solubility	No	Soluble	No	No
Conductivity/solid	No	No	No	Yes
Conductivity/water solution	Yes	No	No	No
Conductivity/ cyclohexane solution	No	No	No	No

With comparable data, students should have little difficulty recognizing basic differences in behavior among different classes of solids. For example, NaCl has a characteristically high melting point, is hard but brittle, is water soluble, and its water solution conducts electricity. By contrast, lauric acid has almost diametrically opposite behavior. It is thus fairly simple to relate properties to the type of bonding during the post-laboratory discussion.

Answers to Implications and Applications

1. Molecules held together by weak van der Waals forces, as in lauric acid, can easily escape from the solid. This is why this substance has an easily detected odor, and also why it deforms easily. The weak intermolecular bonding force explains the low melting point. Some highly volatile substances such as hydrogen and oxygen gas are odorless. The smaller the amount of substance that can be detected by the human nose, the smaller the amount that needs to be present to be detected.
2. Molecules of lauric acid are held in the solid state by weak van der Waals forces, hence it has a low melting point (< 100 °C). The other three substances have strong ionic or covalent bonding forces as solids. NaCl is an ionic solid (m.p. about 800 °C), SiO_2 a covalent network solid (m.p. > 2000 °C), and Fe a metal (m.p. > 2000 °C).
3. No, not really. With lauric acid, weak intermolecular forces are overcome easily without affecting the intramolecular covalent bonds. On the other hand, when NaCl melts, strong ionic interactions must be overcome to separate the ions.
4. The intramolecular forces are strongest in lauric acid. These consist of covalent bonds holding the carbon, hydrogen, and oxygen together in molecules. The intermolecular forces are moderately weak van der Waals forces; consequently the molecules are rather loosely held within the crystal and the crystal lattice is easily disrupted.

5. NaCl dissolves in water but not in cyclohexane, while lauric acid dissolves in cyclohexane but not in water. Neither silica nor iron dissolve. Water consists of polar molecules; their polar nature allows them to overcome the attractive forces in NaCl (and many other polar solutes) sufficiently for the substance to dissolve. Since lauric acid is nonpolar, it does not interact with water strongly enough to dissolve. Cyclohexane is a nonpolar solvent and can dissolve lauric acid. The very strong attractive forces within the other two solids prevent their solution.
6. Student answers will vary depending upon the identity of the unknown. In any case, their answers should be based on the data collected for the unknown and should justify the classification as one of the four classes of substances.

Post-Laboratory Discussion

A systematic discussion on the types of bonding in various solids should take place after students have collected their data. Some generalizations can be drawn:

To a large degree, the physical properties of solids are determined by the type of bonding holding the atoms, molecules or ions in solid form. The bonding gives rise to ionic solids, molecular solids, metals, and covalent network solids; each with its set of characteristic physical properties. Thus, the type of solid can frequently be determined by studying its physical properties.

Ionic solids consist of oppositely charged ions arranged in a crystal lattice in a way that allows each ion to be surrounded by either two, four, six, or eight oppositely charged ions. This causes very strong attraction, giving typically hard solids with high melting points. Ionic solids are frequently soluble in water but insoluble in nonpolar, organic solvents. As solids, they do not conduct electricity, but do so both when fused (melted) or in water solution. However, not all ionic compounds are water soluble, *e.g,* calcium carbonate. In these cases attractions among ions in the lattice are too strong to be overcome by the attraction of polar water molecules.

Molecular solids are composed of molecules held together by relatively weak forces. Thus, these solids tend to be soft, easily melted, and more or less volatile. Molecular solids are likely to be insoluble in water but soluble in organic solvents. Neither the solids nor their melted states conduct electricity, nor do any of their solutions. A few molecular solids such as common table sugar (sucrose), however, are water soluble, and thus must consist of polar molecules.

Covalent network solids are bonded by strong covalent forces into a 3-, 2-, or 1-dimensional network where every atom is linked to every other atom. The "molecules" are macrosized so whatever size sample is being studied can be considered to be a molecule. Actually, the term molecule has little meaning in this regard. The strong multidimensional bonding produces solids with very high melting points that are generally hard, nonvolatile, and insoluble in all solvents. Covalent network solids are nonconductors of electricity with the exception of graphite, a 2-dimensional network solid.

Metal atoms are bonded by metallic bonds. These bonds are usually strong (notable exceptions are mercury, a liquid at room temperature, and gallium, which can melt in a person's hand) but not highly directional (as with covalent bonds). The nondirectional nature of the metallic bond is due to relatively mobile valence electrons and accounts for the malleability of metals. Most metals have 6, 8, or 12 nearest neighbors, but few valence electrons. Although the few valence electrons are free to move about, the atoms are not, since most metals are solids at room temperature and have definite shapes. So we deduce that metal atoms behave as if they are spherical and packed in regular arrays. There are almost flat planes (glide planes) between atom layers. Thus motion along the planes is easy. These characteristics give rise to the good electrical and heat conductivity of many metals as well as their malleability, nonvolatility, and shiny metallic luster. Although metals do not dissolve in water or organic solvents, they do form solid solutions in the form of alloys and amalgams with mercury.

Assessing Laboratory Learning

Student response to identifying an unknown is a good way to assess how much they learned in this activity. Demonstration of important manipulative techniques can also give a clue to students' experimental progress.

It is important that students "see" the relationships among bonding, structure, and properties. This activity reinforces such learning and stimulates numerous questions that can measure student comprehension. For example, it is possible to provide "experimental data" similar to that developed in the laboratory for a particular substance and challenge students to classify the substance as ionic solid, covalent network solid, molecular solid, or metal.

Activity 2: Metal and Ionic Crystals

Introduction

In this investigation, plastic spheres are used to study the regular geometric shapes of metal and ionic crystals. The regular geometric shapes of crystals reflect the orderly arrangement of the atoms, ions, or molecules that make up the crystal lattice.

Three types of packing in crystals will be investigated—hexagonal close packing, face-centered close packing, and body-centered packing—using equivalent spheres. From the models it will be possible to determine the coordination number of the particles of each structure.

The relative sizes of cations and anions that make up ionic crystals are a determining factor for the coordination number. The effect of size as indicated by ionic radius on coordination number will be examined by building models of rock salt (NaCl) and Wurtzite lattices (ZnS).

Purpose

To gain familiarity with the geometry of metallic and ionic crystal structures.

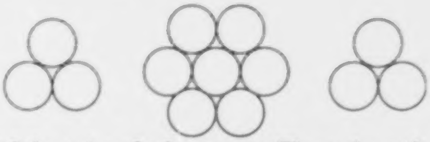
Safety

Wear protective goggles throughout the laboratory activity.

Procedure

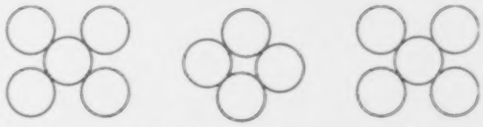
Secure the following per group of two students: 36 two-inch, 13 one-inch, and 13 three-quarter-inch expanded polystyrene spheres. Toothpicks or short lengths of pipestem cleaners can be used as connectors for the spheres.

Part I



1. Use connectors to build these three structures in Figure 1.
2. Place one three-sphere layer on the table or desk so that one apex of the triangle faces you. Then place the seven-sphere layer on top of the three-sphere layer so that the center sphere fits into the depression at the center of the first layer. Now place the second three-sphere layer over the center sphere so the top layer is directly above the bottom layer. This resulting structure represents hexagonal close packing (hcp).
3. Count the number of spheres closest to the central sphere in the structure. This is called the number of nearest neighbors or the coordination number. Zinc, magnesium, and many other metals pack in this manner. Retain this model for Step 5. Answer *Implications and Applications* Question 1 now.

Part II



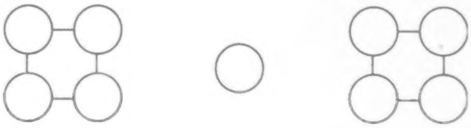
4. Build these layers in figure 2.

   Place one five-sphere layer on the table or desk. Then place the four-sphere layer over the first so that the four spheres rest in the spaces between the corner spheres in the bottom layer. Now place the last five-sphere layer on top so that it is situated directly over the bottom layer. Study the structure carefully and decide why it is called face-centered cubic (fcc). Copper, silver, aluminum, and many other metals pack in this manner in crystals. Determine the coordination number. Answer *Implications and Applications* Question 2.

5. Return to the model of hexagonal close packing built in Part I. Rearrange the layers so that the top three-sphere layer is rotated through 60° and is no longer directly over the bottom layer. The rearrangement changed the type of packing from hexagonal to face-centered cubic. To confirm this, rotate the structure until four spheres forming a square face outward. When this orientation is obtained, remove the top five-sphere layer of the second model (face centered-cubic) and place it over the four spheres facing outward. Note that this results in a fcc packing tilted at an angle to the desk top. Answer *Implications and Applications* Questions 3 and 4.

Part III

6. Use two-inch spheres to build the layers shown in figure 3. The space between spheres in each four-sphere layer should be about 5 mm.



Place the single sphere in the center of one layer and then place the second layer directly over the bottom layer. Study the model carefully to decide about the appropriateness of the name body centered-cubic (bcc). Sodium, potassium and a few other metals pack in this manner. Answer *Implications and Applications* Question 5.

Part IV

7. To construct the sodium chloride lattice, you'll need the structure of Part II (fcc) and 13 one-inch spheres. The two-inch spheres represent the face centered arrangement of chloride ions. Insert the 13 one-inch spheres (representing sodium ions) into the holes between the chloride ions in each layer. Note that there is also a face-centered arrangement of the one-inch (sodium) spheres.
8. Determine the coordination number of each size sphere. The sodium ion has a diameter of 0.19 nanometers; the chloride ion's diameter is 0.36 nm. This gives a ratio of about 1/2, so one-inch and two-inch spheres give a good approximation of the NaCl lattice. Answer *Implications and Applications* Questions 6 and 7.

Part V

9. Zinc ions have a diameter of 0.15 nm while sulfide ions have a diameter of 0.37 nm. Use two-inch spheres to represent sulfide ions and three-quarter-inch spheres to represent zinc ions. Use the Part I structure (hcp) to represent sulfide ions. Using connectors, secure a three-fourth-inch sphere above each of the two-inch spheres in each of the three hcp structure layers. Then place the largest layer on the table and fit one of the smaller layers on top so the smaller spheres fit into alternate depressions. Invert these two layers and fit the other small layer, small spheres pointing upward, above the larger layer so the top-layer spheres are directly above the spheres on the bottom layer. Study the structure carefully. Answer *Implications and Applications* Question 8.
10. Return the spheres and connectors after dismantling the models.
11. Thoroughly wash your hands before leaving the laboratory.

Implications and Applications

1. What is the coordination number for the center atom in the hexagonal close packing model?
2. Why is the name face-centered cubic appropriate for this structure? Explain.
3. Is the coordination number for the two types of close-packing of equivalent spheres different?
4. How would the density of hexagonal close packing and face-centered cubic compare if both were constructed with spheres of the same size and mass?
5. Metallic iron crystallizes in a body-centered cubic form below 906 °C. It is called alpha-ferrite. Above this temperature, iron forms face-centered cubic crystals, called gamma-ferrite. When iron cools to 140 °C, it changes back to body-centered cubic called delta-ferrite. What is the coordination number of each form of iron?
6. Which ions most closely surround each Cl^- ion? Which ions most closely surround each Na^+ ion?
7. What is the coordination number of the spheres representing Na^+ ions? What is the coordination number of the spheres representing Cl^- ions?
8. What is the coordination number of the spheres representing both the Zn^{2+} and the S^{2-} ions?

Activity 2 was adapted from Experiment 27, pp. 72-74, in the following source:

Malm, L. E. (Ed.). (1963). *Chemistry: An experimental science. Laboratory manual.* San Francisco, CA: W. H. Freeman and Company.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Metal and Ionic Crystals

Major Chemical Concept

This laboratory activity gives students an opportunity to view the basic structural forms of most metals. Students learn the meaning of *coordination number*, *crystal lattice*, and *packing*. Later, students will be able to correlate properties with structure.

Two common ionic crystal lattices—rock salt and wurtzite—and their coordination numbers give students some insight into why ionic solids are brittle with relatively high melting points, and why different-sized ions form different types of crystal lattices. Each ion is surrounded by oppositely charged ions, producing strong attractions, holding the ions in place.

Level

This activity is appropriate for all levels of students. Basic students probably should not complete the extensions.

Expected Student Background

Students with no previous knowledge of crystals can perform the activity. There are no prerequisite concepts other than those listed for this topic. The activity could also be used to support the topic of condensed states of matter.

Time

Students should be able to complete the activity in one 50-min period.

Safety

No special precautions are necessary. If students work in the chemistry laboratory, they should wear protective goggles since there may be hazardous materials nearby. Advise students to follow regular laboratory safety procedures.

Materials (For 24 students working in pairs)

432 Polystyrene spheres (available from hobby shops), 2-inch
156 Polystyrene spheres, 1-inch
156 Polystyrene spheres, 3/4-inch
Pipestem cleaner lengths, 2-cm, or toothpicks

Advance Preparation

Have sufficient spheres available and placed either in boxes for each two-student group or readily obtainable by students. Consider preparing a set of layers for the models glued together to show students what the layers should look like.

Pre-Laboratory Discussion

Minimal pre-laboratory discussion is required. If glued-together layers are available, they could be shown to students. Directions for connecting spheres should be given and also demonstrated.

Teacher-Student Interaction

Students will probably need help in determining coordination number, particularly with the wurtzite structure. If a set of layers that students construct are permanently glued together prior to the activity, the task of helping students with coordination number, layer orientation, *etc.*, is greatly simplified.

Visit with students about the meaning of metallic and ionic bonding, and regarding the question on the density of two models of close packing of equivalent spheres. Considerable individual teaching can take place through such interactions. Students may also need help answering some of the other questions. Encourage students to record observations and answer questions as they perform the activity.

Anticipated Student Results

It is not necessary to have data tables. Students may record observations directly on their laboratory sheets and answer the questions immediately. Most students can do the model building without difficulty. Some may have difficulty with coordination number until they have the concept demonstrated for them.

Answers to Implications and Applications

1. The coordination number is 12; the central atom has 12 nearest neighbors.
2. Face-centered cubic packing has a sphere in the center of each cube face, hence the name.
3. No. The two types of close packing both have a coordination number of 12.
4. The density would be the same since both use the same volume with the same efficiency.
5. Alpha-ferrite is body centered cubic with a coordination number of 8. Gamma-ferrite is face centered cubic with a coordination number of 12. Delta-ferrite is body centered cubic with a coordination number of 8.
6. Each sodium ion is most closely surrounded by chloride ions and each chloride ion is most closely surrounded by sodium ions.
7. The coordination number of each sodium ion is 6, which is also the coordination number of each chloride ion.
8. The coordination number of sulfide ion is 4; it is also the coordination number of zinc ion in the Wurtzite lattice.

Post-Laboratory Activities

Discuss student answers to the questions. Have models available to show students any points they might have missed. Use the laboratory results in class discussion concerning metallic bonding and ionic bonding.

Extensions

1. Build models from two-inch spheres with tetrahedral (four spheres) and octahedral (six spheres) structures. Determine which size spheres—three-fourth-inch or one-inch—will fit the "holes" in these structures.
2. Calculate the ratios of smaller-to-larger spheres in each case (0.75/2 and 1/2). Calculate the diameter ratios of Zn^{2+}/S^{2-} and Na^{+}/Cl^{-} (Zn^{2+} diameter = 0.15 nm, S^{2-} diameter = 0.37 nm, Na^{+} diameter = 0.19 nm, Cl^{-} diameter = 0.36 nm). Are the two-inch, one-inch, and three-fourth-inch spheres good models for these ions?

Assessing Laboratory Learning

A good way to assess student learning in this activity is to check student laboratory reports. A good time to check for understanding is while students are performing the activity. If they can answer the questions asked, count nearest neighbors, and build the models, then they have demonstrated good understanding of the material. Here are two possible student questions:

1. Explain how face-centered, hexagonal close packed, and body-centered cubic crystals are different.
2. The chemically similar alkali metal chlorides NaCl and CsCl have different crystal structures (fcc and bcc) while the chemically different NaCl and MnS have the same crystal structures (fcc). Why?

DEMONSTRATIONS

Demonstration 1: Electrical Conductivity of Solutions

Purpose

This demonstration provides experimental evidence on the nature of ionic and molecular substances in solution and as a fused (molten) ionic solid. The major purpose is to show that ionic solids conduct the electric current both in solution and when fused, whereas molecular solids do not.

Materials

Electrical conductivity apparatus (commercial or home-made)
6-8 Beakers, 50-mL, or 2-oz. wide-mouth bottles
Solid sodium chloride, NaCl, and sucrose, $C_{12}H_{22}O_{11}$ (common table sugar)
Solid silver nitrate, $AgNO_3$
2 or 3 Small crucibles
Ringstand, ring, triangle
0.1 M Solutions of NaCl and sucrose (5.6 g NaCl/100 mL solution and 3.4 g sucrose/100 mL solution)

Safety

If the conductivity tester used is powered by a 120-V source, then use caution to prevent electric shock and DO NOT ALLOW STUDENTS TO USE THE APPARATUS. The substances used are innocuous and do not require special handling except for the silver nitrate. This material is both toxic and caustic. Handle it with care, particularly when it is fused since it is also an oxidizer. Solutions may be safely disposed of by flushing down the drain with water. The solid NaCl and sugar may be used for all demonstrations and then disposed of in the trash. Keep the crucible with silver nitrate in a dark bottle for later use or for next year.

Procedure

Half fill a 50-mL beaker with solid NaCl and a second 50-mL beaker with solid sucrose. Test the electrical conductivity of the solid NaCl with the tester (light bulb remains dark). Clean the electrodes and then test solid sucrose (light bulb remains dark). Have students tell what is observed in each case.

Half fill two 50-mL beakers with the two solutions and a third with distilled water. First test the electrical conductivity of the distilled water (light bulb remains dark). Then test the electrical conductivity of each of the two solutions in turn (NaCl causes light bulb to glow, while bulb remains dark in the sucrose solution). Have students report what is observed in each case.

Half fill a small crucible with silver nitrate. Place the crucible on a triangle supported on an iron ring mounted on a ring stand. Heat the crucible with a laboratory burner until the silver nitrate is melted. Carefully test the electrical conductivity of the melted silver nitrate (light bulb glows brightly). Have students report what they observe. *NOTE: Sodium chloride is used in solution because all students are familiar with it. However, silver nitrate is used as a melt because it melts at a considerably lower temperature than does sodium chloride.*

Remarks

Approximately 30 min are needed to prepare the demonstration and to "dry run" it. From 10 to 20 min (depending on teacher input and student interaction) are needed to perform the demonstration.

The demonstration may be used either in the pre-laboratory discussion of either of the laboratory activities, during the post-laboratory discussion, or it may be used to illustrate properties of ionic solids and molecular solids. In any case it illustrates an important property of ionic solids—electrical conductivity—and is worth doing. An expanded form of this demonstration may also be used in the *Condensed States* module.

A suitable home-made electrical conductivity tester can be simple or fancy*. A simple one consists of two porcelain light bulb receptacles mounted base-to-base on a ping pong paddle. These are wired in series with an electrical cord fitted with an on-off switch. Place the base with intact heavy wires of a 150-W light bulb that has the glass envelope removed in one receptacle and a 25-W clear glass light bulb in the other. The two wires in the 150-W base act as the test electrodes. To avoid electrical shock, always turn off the switch or pull the plug while cleaning or rinsing the electrodes or changing the solutions.

Materials may be safely disposed of by flushing solutions down the drain with water, and, after using the solids for all classes, they may be placed with solid waste. Due to the expense involved, keep the crucible with silver nitrate in a dark bottle for use later or next year.

Discussion

The results can be conveniently used to define an electrolyte and a nonelectrolyte (the former exists as ions in water solution while the latter does not). The solids do not conduct since the ions in sodium chloride are not mobile and cannot move toward the electrodes and act as charge carriers. Sucrose is a molecular solid and does not have ions present in either the solid or solution. The melted silver nitrate conducts because the ions are able to move to the electrodes and act as charge carriers. This emphasizes that an ionic solid consists of ions and not molecules.

Do not get involved in a discussion of electrical conductivity now, since the students' backgrounds are probably insufficient for understanding. The point is to distinguish between electrolytes and nonelectrolytes and to show that ionic solids exist as ions in the solid state. If students ask about the lack of conductivity of distilled water, you can test the conductivity of tap water. Tap water conducts because there are ions present in it but not in the distilled water.

Based in part on the demonstration "Electrical Conductivity" in:

McClellan, A. L. (Ed.). (1963). *Teacher's guide for chemistry: An experimental science* (pp. 157-158). San Francisco, CA: W. H. Freeman and Company.

*References for making a low-cost, safe apparatus for conductivity testing may be found in the *Acids and Bases* module in the *Instrumentation* section.

Demonstration 2: Displaying Structures in Two Dimensions

Purpose

To reinforce the usual pencil-paper (or chalkboard) presentation of Lewis-dot formulas

Materials

Homemade or commercial flannel board (about 2 x 3-feet)
30-40 Circles, 1-in diameter, of tag board or light cardboard with a felt backing
10-20 Circles, 3-in diameter, of tag board or light cardboard with a felt backing
Overhead projector and screen
30 Circles, about 1-in diameter, cut from colored transparency film
30 Circles, about 1/4-in diameter, cut from colored transparency film
(These are easily produced from colored transparency film and an ordinary paper hole punch.)
Transparency grid with 1 cm x 1 cm rulings

Procedure

Demonstrate the Lewis-dot formulas for several atoms, molecules and ions—for example, C, O, N, Ne, H, Al^{3+}, K^+, CH_4, H_2O, NH_3, Cl_2O, H_2S, *etc.* Limit examples to binary (two-element) molecules if possible. Write the same structures on the chalkboard while demonstrating to reinforce the concept and to show students how to use paper and pencil to do this.

After your demonstration, consider asking a student to come to the flannel board or overhead projector to try to illustrate the bonding and nonbonding pairs for a particular molecule. If time permits, have other students do the same.

Hand out the duplicated construction paper to students and encourage them to practice at home. Provide a list of elements, molecules, and ions for them to use.

Remarks

This demonstration may be used to introduce the concept of Lewis-dot structures along with a chalkboard presentation. It requires about 90 min to prepare the flannel board and cutouts. Presentation time may vary from 5 to 50 min depending upon how much is presented and how much discussion is generated.

Use small circles to represent electrons and large circles to represent atoms or ions. If colored transparency stock is available along with an overhead projector, then it is probably easier and simpler to use this method. This will also cut down on storage requirements.

Students can be furnished with circles of two diameters (1-in and 1/4-in are good) duplicated on colored construction paper to take home and cut out for practice with Lewis-dot formulas.

While showing students the Lewis-dot structures, point out steps involved in writing Lewis-dot formulas. State and apply the octet rule as the demonstration progresses. If students are called on, have them show that the octet rule is followed.

This is a good time to introduce the use of double and triple bonds to show the octet rule and common patterns when C, N, O, and S are involved. For example, CO_2 and CO illustrate this nicely.

The use of physical models such as these to represent atoms, and, especially bonds, is only an approximation of what chemists think is "real." [Based on Demonstration B23, Displaying structures in two dimensions. (1988). *Doing Chemistry*. Washington, DC: American Chemical Society.]

Demonstration 3: Magnetic Analogy for Bonding Forces

Purpose

To demonstrate the electrostatic attraction and repulsion between like and unlike charges using a magnetic analog.

Materials

4 Ceramic ring or disk magnets
Overhead projector and screen

Procedure

Begin the demonstration by showing on the overhead projector that like poles repel and unlike poles attract. Draw an analogy to attraction and repulsion between electric charges.

Set up two pairs of magnets so that each pair is attractive. This will illustrate the attraction between the nucleus and an electron. Move one pair into the center of view. Then approach with the other pair in so that one "nucleus" approaches the other "nucleus." No attraction is observed. Move this pair so that the "electrons" are between the two "nuclei." This leads to a stable arrangement.

With the stable arrangement in the center, show that if either "electrons" or "nuclei" are forced closer together, they repel. Thus an "equilibrium distance" between particles with like charges is created.

Remarks

This demonstration takes 15 min or less to prepare and present, depending upon availability of disk or circular magnets. The demonstration should be presented when covalent bonding is introduced.

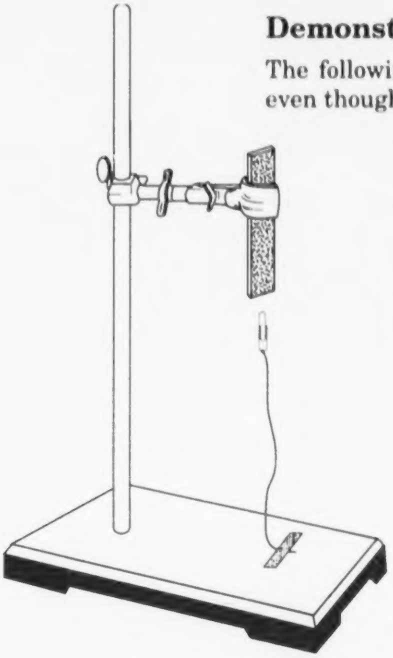
Practice with the magnets prior to doing the demonstration. It requires patience to move the magnets carefully to show formation of a "bond". Either ring or disk magnets may be used, but they should be face-magnetized, that is, faces should be the poles. It is possible to use ring magnets to represent the nuclei and disk magnets to represent the electrons. In either case, use larger-diameter magnets for nuclei and smaller-diameter for electrons.

The materials may be saved and used many times. Ceramic magnets are somewhat fragile and should not be dropped. Edmund Scientific, Radio Shack, and similar sources can supply these magnets. Disk magnets are also frequently available at hobby shops and craft shops.

Point out that this is an analogy to the electrostatic forces that cause a covalent bond to form. Both magnetic and electrostatic forces behave the same way; like charges or poles repel and unlike ones attract. Point out that a "nucleus-electron" pair is stable because there is only one attractive force and no repulsive forces. When two pairs approach, new attractive forces arise between the "nucleus" of one pair and the "electron" of the other pair, and *vice versa*. New repulsions also are present between the two "nuclei" and the two "electrons." The result is four attractions but only two repulsions, hence the two pairs form a stable arrangement.

Demonstration 4: Force at a Distance

The following demonstration shows that forces can exist even though objects actually are not in direct contact.



A paper clip can be suspended about two millimeters from the magnet. The thread is attached to the base of the ring stand with adhesive tape. A piece of paper may be passed between the paper clip and the magnet without disturbing the setup. A long hair may be substituted for the thread.

Figure 4. Force between a magnet and a paper clip.

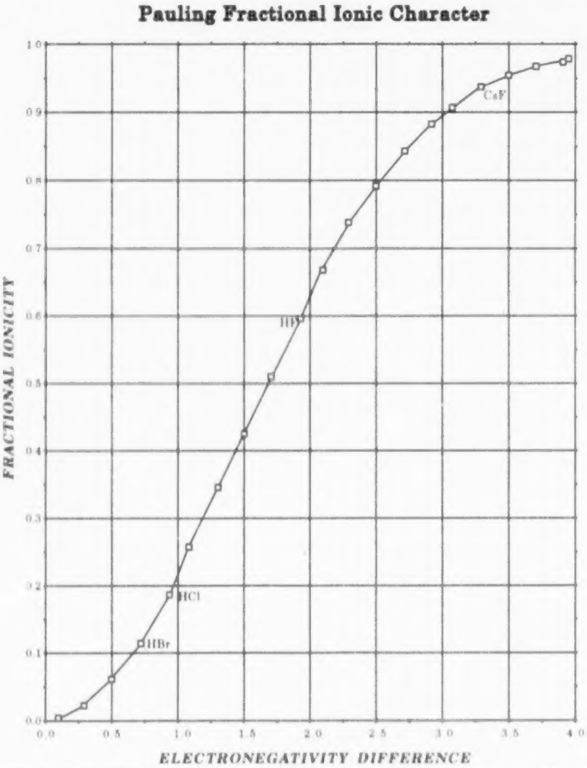
GROUP AND DISCUSSION ACTIVITIES

Counterintuitive Examples, Discrepant Events

No Compound is 100% Ionic: A Group Activity

Several different equations can be used to estimate the degree to which a chemical bond between two atoms is ionic or covalent. The electronegativity difference between the atoms, $(x_A - x_B)$ is a central key. x_A is the electronegativity of the more electronegative element. One of these equations was developed by the famous chemist, Linus Pauling. When this equation is plotted, it has the shape of a stretched-out S. Such a graph is shown in Figure 5 where fractional ionicity represents the fractional ionic character (a value of 1 indicates 100% ionic).

Figure 5. Pauling Fractional Ionic Character. *(See* Appendix*)*



Procedure

1. Obtain a graph of fractional ionicity versus electronegativity difference.
2. Look up electronegativities for the atoms in these six compounds: KBr, H_2S, H_2O, HI, RbF, and NaCl. Calculate six values for $x_A - x_B$.
3. Locate the positions of these compounds on the graph. Mark each compound on the graph.
4. Estimate from the curve the value for fractional ionicity when $x_A - x_B = 4$.
5. Estimate from the curve the value for fractional ionicity when $x_A - x_B = 0.5$, 1.0, 1.5, 2.0, 2.5, 3.0, and 3.5.

   (*Teacher's Note*: Many textbooks refer to $x_A - x_B = 1.7$ [50% ionicity] as an arbitrary demarcation between covalent and ionic bonds. Reference to the graph shows the futility of trying to find a "boundary" between the two types. For example, according to this arbitrary rule, HF would be ionic; however, experiment shows this compound is polar covalent. Other examples underscore the fact that there is really no way to classify most compounds as being either ionic or covalent. There is a gradual increase in ionicity when electronegativity difference increases, as shown by the graph. Thus, it is reasonable to state that when ionicity is greater than 50%, there is considerable ionic character to the bond. Typical ionic properties are usually encountered at greater values of ionicity. However, there is only one kind of "pure" bond, the homonuclear covalent bond.

What does all of this show?

1. An electron is never *completely* transferred from one atom to another to form an ionic compound.
2. An electron pair is always shared to some degree between two atoms.
3. The more equal the sharing, the higher the covalent nature of that compound.
4. **Rule of Thumb**: If a compound has an ionicity greater than 50%, (some authorities use an electronegativity difference of 2.0 representing about 63% ionicity) the compound is defined as "ionic." So, to simplify life, an arbitrary classification "ionic" is used.

At this point, it might be well to define ionic solid operationally as one whose water solution is a good conductor of electric current. If a compound is insoluble in water, and if in a fused (melted) form conducts electric current, then the compound may be classified as ionic.

Analogies and Metaphors

Resonance

The actual form of a molecule having resonance structures is like a central party in politics, a moderate denomination in religion, the liquid state of matter, or mixing red and blue liquids.

1. The extremes define and "bracket" an intermediate concept, theory, or idea. Frequently, something is described by saying what it is *not* like, *e.g.*, "Dan, I'm glad I'm not *that* narrow-minded." The terms liberal Republican, conservative Democrat, and orthodox Jew, for example, are used to convey deviation from, or strict adherence to, an idea or concept. The British Social Democratic Party was founded to offer a centrist alternative to more extreme positions of the ruling Conservatives on the right and the Labour Party on the left.

Likewise, the liquid state lies between the solid state and the gaseous state in the sense that it is less dense than the solid state and more dense than the gaseous state. Solids and gases are extremes. Solids are usually very ordered in the arrangement of their atoms and molecules, whereas gases show no ordered arrangement of their atoms and molecules. Liquids have a "short-range order," but no "long-range order." So the liquid state has been defined by what it is *not* like.

2. Orthodox (or fundamental) religion, and right and left wing extremism in politics are common terms. Orthodoxy and fundamentalism imply strict adherence to an idea or a concept. A specific chemical "resonance structure" is "orthodox"—it is exactly that structure and nothing else! A different resonance structure of the same molecule is that extreme as written and nothing else. However, the concept to be conveyed by the idea of resonance is that the actual molecule for which resonance is possible is none and yet all of the extreme or orthodox forms. It is an intermediate among resonance forms that we are able to define and draw.
3. When blue and red colors are mixed (*e.g.*, dye solutions) the resulting color is neither blue nor red, but something intermediate. In other words, a "resonance hybrid" of the two individual colors. This might be illustrated as follows:

   Blue + Red → Purple
4. The picture of "extreme religions" or political parties as developed in the context of the chemical concept of resonance furnishes a good lead into a discussion of resonance. Remember, the "resonance" concept implies the existence of many facets, faces, and forms, but the real picture is none of those extremes entirely!

Covalent Bonding

The covalent bond illustrated by the two friends sharing an ice cream cone is like a single covalent bond, sometimes called a sigma (σ) bond. If, while sharing the cone, they also share the burden of carrying a canoe overhead, this would be like a double covalent bond, sometimes called a pi (π) bond.

"Bonding" Contexts

Bonding is a term commonly used in contexts other than in chemistry. Discuss the following statements as common examples of the use of the term bonding.

1. "I'm really bonded to the Rocky Mountains because I was born in Colorado Springs."
2. "I'm really bonded to the Texas Hill country because my grandmother's house was only fifty miles from Austin."
3. The accused house burglar had to post a $5000 bond to appear in court.
4. "My uncle gave me a $50 U.S. Savings Bond when I was 10 years old."

Polarity of Molecules

Teflon-coated stirring bar magnets on an overhead projector can be used to illustrate an analogy between magnet behavior and polar molecule behavior. If the stirrers are spaced away from each other with a random orientation, then when a magnet passes near them on the overhead light stage, the magnetic stirrers will torque to align themselves parallel to the magnet's magnetic field. Polar covalent molecules behave similarly in an electric field.

Key Questions

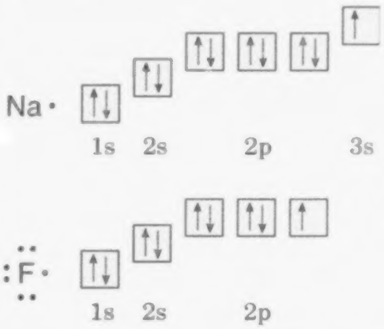
1. What are valence electrons? *[Electrons in the highest occupied energy level of an atom are known as valence electrons. These electrons determine what kind of chemical bonds, if any, the atom can form.]*

2. How can the total valence electrons for an element be determined? Explain. *[Electron configurations may be used to determine the number of valence electrons for an element. For example, hydrogen has only one electron—its valence electron. The configuration is* $1s^1$*. For representative elements, the number of valence electrons equals the total electron population at the highest principal energy level (n), as indicated by electron configurations.]*

3. How many valence electrons does each type of atom have: sodium, silicon, beryllium, and oxygen? *[sodium = 1; silicon = 4; beryllium = 2; oxygen = 6]*

4. What is the relation between the number of valence electrons in atoms of an element and the element's placement in the periodic table? Give examples. *[The number of valence electrons determines the group placement of an element. For example, hydrogen has one valence electron; it is in the alkali metal family. All other elements in this family, Li, Na, K, Rb, and Cs, also have only one valence electron. On the other hand, fluorine has seven valence electrons, as shown by its configuration* $1s^22s^22p^5$*. This places it in the halogen family. (NOTE: This is probably more than the average student will give as an answer.)]*

5. Draw orbital diagrams for atoms of sodium and fluorine. Use the diagrams to write Lewis-dot formulas for these elements.

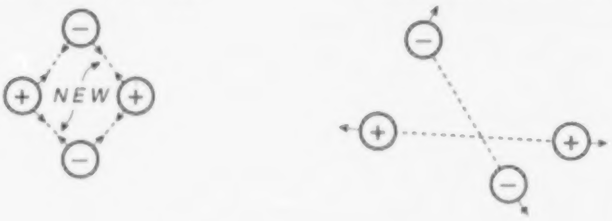


6. Discuss how the Periodic Table helps to determine the number of valence electrons for an element. Include examples and state the trend in valence electrons within a group of the Periodic Table. *[An alkali metal group element has one valence electron, alkaline earth elements have two, boron group has 3, and so forth. Valence electrons determine chemical properties that, in turn, determine element placements in the table. For example, fluorine has seven valence electrons; all other halogens—Cl, Br, and I—likewise have seven valence electrons. (Actual student answers may vary considerably.)]*

7. Why are molecules more stable than separated atoms, particularly among representative elements? *[Student responses will vary from simple to complex. The following answer is more sophisticated than that expected of an average student: The origin of chemical bond stability depends upon the attractive and repulsive electrostatic forces present. Electron-nucleus interaction always furnishes attractive forces while nucleus-nucleus and electron-electron interactions furnish repulsive forces. A simple way to show the origin of this stability is in terms of possible attraction for one or more electrons simultaneously by two nuclei close to each other. Whether two given atoms can form a bond depends upon the filling of orbitals in the separate atoms. Figure 6 illustrates this concept. The hydrogen molecule is used to show that when two hydrogen atoms are close together, there is a possibility of more attractive forces than repulsive forces. On the other hand, when two helium atoms approach, a net attractive force does not occur, although there are four new attractive forces. This can be explained—at least in terms of this simple analysis—by the equal number of new repulsive forces also formed, as shown in Figure 6.]*

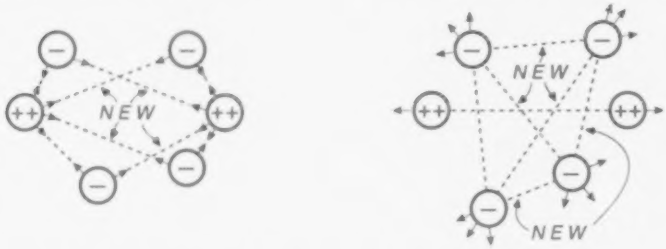
The Hydrogen Molecule

Attractive Forces *Repulsive Forces*



Approaching Helium Atoms

Attractive Forces *Repulsive Forces*



This is similar to transparency masters 1 and 2.

Figure 6. Bonding of the hydrogen molecule.

8. How many bonds may be formed by atoms of representative elements? Explain, including all groups of representative elements. *[Most students will probably give the number of bonds for each periodic group they have memorized, correlating the number of bonds with the number of valence electrons. The answer given here is more detailed than what typically good students would provide: The number of bonds that an atom of a representative element can form depends upon the orbital occupancy of valence electrons in the atom. Beginning with the halogen family, the valence electron shell has seven electrons, three pairs and one unpaired electron in s- and p-orbitals—* $3s^2 3p_x^2 3p_y^2 3p_z^1$ *in chlorine, for example. This gives any halogen elements the possibility of forming a single covalent bond by forming one electron pair with an unpaired valence electron from another atom. This is shown by the formation of HF, illustrated in Figure 7.*

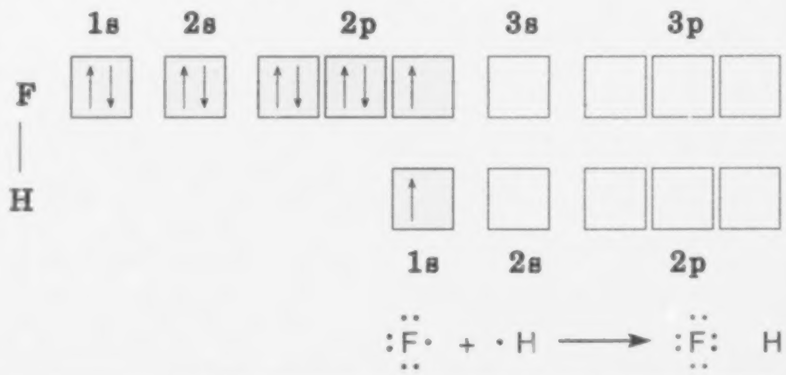


Figure 7. The bonding of HF.

In the case of oxygen family elements (also called chalcogens), each element's atom has six valence electrons; two of these are unpaired—$2s^2 2p_x^{\,2} 2p_y^{\,1} 2p_z^{\,1}$ in oxygen, for example. Two single covalent bonds, as for oxygen in water, may be formed under these circumstances, as illustrated in Figure 8.

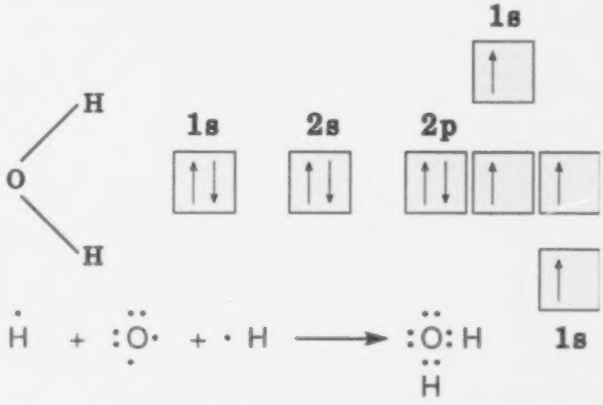


Figure 8. The bonding of H_2O.

A similar analysis for nitrogen family elements gives three unpaired valence electrons, one in each valence shell p-orbital, providing the possibility of three covalent bonds, as in the formation of NH_3 illustrated in Figure 9.

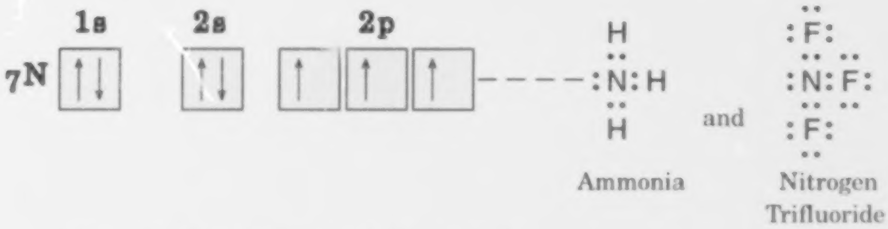


Figure 9. The bonding of NH_3 and NF_3.

With carbon, the situation is somewhat different. Carbon has the valence configuration $2s^2 2p_x^{\,1} 2p_y^{\,1} 2p_z^{\,0}$ indicating that it can form two covalent bonds. However, almost all carbon compounds form four, not two, covalent bonds. Two new ideas come into play here. First, the empty valence orbital, if CH_2 formed, would give it chemical activity (reactivity) due to its close proximity to the nucleus. Second, the four equivalent bonds formed by many carbon compounds—CH_4 for example—can be explained by imagining that one 2s electron is moved to the empty 2p orbital. This process, called promotion, requires energy input in a hypothetical intermediate step that is not needed in any previous cases. The energy difference between the 2s and the 2p orbitals is small, however, and does not prevent formation of four equivalent bonds. Methane, CH_4, illustrates this (see Figure 10).

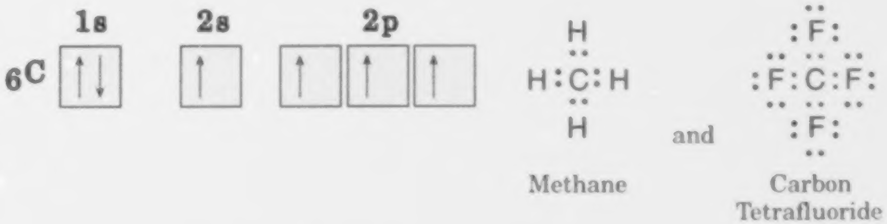


Figure 10. The bonding of CH_4 and CF_4.

A more rigorous consideration views promotion as a step in the hybridization of atomic orbitals to form equivalent bonding orbitals. Thus, in CH_4, carbon's s-orbital and three p-orbitals are combined mathematically to form four equivalent hybrid orbitals that are used when carbon bonds with four hydrogen atoms. These four hybrid orbitals contain both s and p-orbital characteristics and achieve maximum separation by pointing toward the vertices of a regular tetrahedron at angles of 109° 29′. This is designated as sp^3 (s-p-three) hybridization since one s- and three p-orbitals in carbon are involved in forming four equivalent hybrid orbitals. Each hybrid orbital, like each original atomic orbital, has a two-electron capacity. NOTE: Similar hybrid orbital arguments can be used to explain the geometries of the two previous examples. NH_3 is a triagonal pyramid; H_2O is a bent molecule (see Molecular Geometry *module).*

All ideas needed to explain the bonding of remaining molecule-forming representative elements have now been described. In the case of boron, valence configuration $2s^2 2p_x{}^1 2p_y{}^0 2p_z{}^0$, promoting one of the 2s electrons allows formation of three equivalent bonds as in BCl_3. Orbital diagrams and Lewis-dot formulas similar to those in Figures 7-10 can be drawn. In terms of hybrid bonding, three equivalent bonding orbitals designated sp^2 (s-p-two) are formed.

Beryllium reacts with hydrogen to form BeH_2. Using similar reasoning to that used above, this bonding can be regarded as involving of two equivalent hybrid bonding orbitals designated as sp hybridization.

Finally, lithium reacts with hydrogen to form LiH.]

9. What is the octet rule? *[When atoms react, they often change electron populations to acquire the stable electron configuration of a noble gas—eight electrons in the outer energy level. For hydrogen, of course, this would be a "duet" rule with two hydrogen atoms sharing two electrons. (This is not a hard rule since there are many exceptions. It is useful in predicting the bonding expected when many atoms form compounds. If the term "octet rule" is objectionable, an alternative is to point out that atoms tend to seek noble gas electron configurations either by electron sharing (covalent bonding) or transferring (ionic bonding) when forming compounds.]*

10. In chemical reactions, do metals and nonmetals behave the same or differently with respect to sharing or transferring electrons? *[Metals generally have lower electronegativities than do nonmetals. Thus, metal atoms attract electrons less strongly and tend to lose electrons to acquire an octet (noble gas electron configuration). This gives the metal atom a net positive charge, resulting in a cation. Nonmetals, on the other hand, behave in the opposite manner, having higher electronegativities than metals. Nonmetal atoms tend to gain electrons to acquire a noble gas electron configuration, giving them a net negative electric charge. Nonmetals tend to gain electrons and form negatively charged ions (anions). These generalities hold reasonably well for many reactions involving representative elements, particularly if higher members of the carbon family are excluded.]*

11. Draw orbital diagrams for the sodium ion, Na^+, and the chloride ion, Cl^-, showing the outermost energy level only. Then write the Lewis-dot formulas and electron configurations for these species.

(TEACHER'S NOTE: Any simple monatomic ion will work with this question, e.g., *Li^+, F^-, Mg^{2+}, S^{2-},* etc. *Lewis-dot formulas of single ions are usually enclosed in brackets and the ionic charge indicated.)*

Na^+	↑↓	↑↓	↑↓	↑↓	$[Na\ ]^+$	$2s^2 2p^6$
	2s		2p			
Cl^-	↑↓	↑↓	↑↓	↑↓	$[:\overset{\cdot\cdot}{\underset{\cdot\cdot}{Cl}}:]^-$	$3s^2 3p^6$
	3s		3p			

12. Draw Lewis-dot structures for NH_3, H_2O, Cl_2O, C_2H_4, and SiO_2. *Many compounds may be used for this question. Avoid "problem" molecules such as those with an odd number of valence electrons that cannot follow the octet rule (such as NO), at least initially. When teaching students how to draw Lewis-dot structures, a useful technique is to develop a helpful set of rules, such as these:*

    a. Count the total number of valence electrons in the structure.

    b. For ions, add one electron for each negative charge and subtract one electron for each positive charge.

    c. Draw the skeleton using dashes to represent electron pairs joining two atoms together until the skeleton is complete.

    d. Add dot-pairs until all valence electrons are accounted for and each atom has an octet of electrons (duet for hydrogen).

    e. If Step d is impossible when N, C, O, or S are involved, try double or triple bonds (two pairs or three pairs of dots) to form octets.

    A simple, quick method for drawing Lewis-dot structures is presented in Tips for the Teacher *under problem solving. It offers as an alternative to the rules given above or for when students have developed some facility drawing these structures.*

13. Compare the covalent bonds formed between elements of similar electronegativity such as carbon and hydrogen and covalent bonds formed between elements with significant difference in electronegativity, such as hydrogen and chlorine. *[Student answers will probably vary considerably. At minimum, some reference to equal or unequal electron sharing should be made. Students should note that unequal sharing produces a charge separation. In any case, students should point out that polar covalent bonds, depending on molecular geometry, often give a molecule properties that affect its behavior.]*

14. Describe the bonding trend expected when fluorine bonds with each element in the second row of the Periodic Table, including itself—F_2, OF_2, NF_3, CF_4, BF_3, BeF_2, and LiF. *[Student answers will vary. At minimum, students should recognize that a variation in bond type from homonuclear covalent to essentially ionic takes place. In their answers, students may note the arbitrary nature of deciding at which point polar covalent bonds are better regarded as ionic.]*

15. What happens to the system's total potential energy when two isolated atoms capable of bonding come into close proximity? *[As atoms approach, atomic nuclei are attracted to valence electrons. If both atoms have half-filled or empty valence orbitals, then bonding may occur, lowering the potential energy of the system due to attractive forces reducing the separation between atoms. See Figure 11.]*

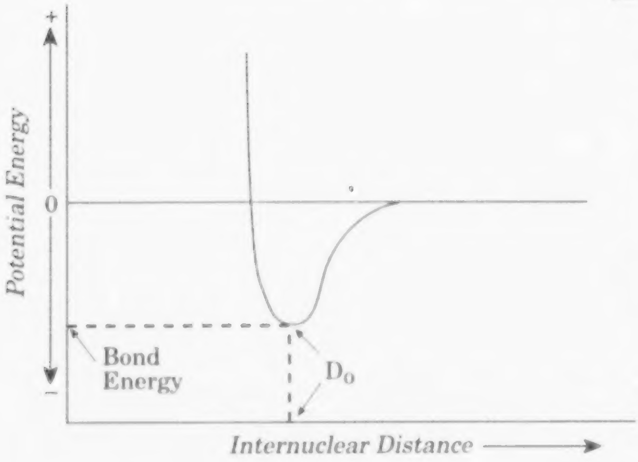


Figure 11. Potential Energy Diagram.

16. Metals have characteristic properties that can be explained in terms of bonding. Briefly describe the bonding in metals that explains such properties. *[Student answers will vary. The basic ideas are found in the post-laboratory discussion for Laboratory Activity 1 involving chemical bonding in solids.]*

17. Some substances are molecular with low melting points, soft structure, and low solubility in water. What kind of bonding could account for these properties? *[Student answers will vary. This is a very open-ended question; it is probable that students will not give complete answers without help from you. Basically, three types of intermolecular interactions—not including the strong covalent bonding within the molecules themselves—may be responsible. These interactions are collectively referred to as van der Waals forces. The weakest intermolecular forces are London forces, as seen in noble gases, carbon dioxide, and other nonpolar, low-melting substances. These forces can be viewed—at least at a basic level—as due to motion of electrons and formation of temporary dipoles. The forces become stronger as the total number of electrons within molecules and the surface area of the molecules increase. For example, carbon dioxide melts at a considerably higher temperature than does molecular hydrogen.*

    Another type of weak intermolecular bonding force is a dipole-dipole force. Two kinds of these forces, distinguished by the energy required to break them, are possible. One is the attractive force between opposite charges of polar molecules. It is stronger than London forces but much weaker than ionic forces in ionic solids. Molecular solids and liquids with this type of bonding generally have higher melting points.

    Some compounds appear to have abnormally high melting points when compared to compounds of similar size, shape, and total electrons. In many of these cases, such compounds exhibit hydrogen bonding, involving highly electronegative nitrogen, oxygen, or fluorine atoms. The small size and high electronegativity of such atoms cause highly unequal sharing of the electron pair forming the covalent bond to hydrogen; there is substantial separation of charge. The intermolecular attractions (between the hydrogen atom and a lone pair of electrons on a N, O, or F atom from an adjacent molecule) arising in this fashion are about an order of magnitude stronger than ordinary dipole-dipole bonding. Water is an example of a hydrogen-bonded substance. If its intermolecular forces were simple dipole-dipole forces, it would melt and boil at lower temperatures than hydrogen sulfide, H_2S, a gas at room temperature.]

Pictures in the Mind

1. **Close Encounters—Sticky Collisions: A Group Activity**

    Have six-student groups, wearing placards hanging from their necks, act out the following scenario while you describe the action. After one run-through, have the entire class practice (in groups of six students) to get a feeling for bonding. This can also be reinforced using appropriate cut-outs on a flannel board or on the overhead projector. Molecules form and chemical reactions can occur when atoms "collide." There are three important types of collisions. These are:

    Close Encounters of the First Kind: Sighting

    An interaction between two atoms in which nothing happens—not a reactive collision—is often called "scattering." (Students acting this out should be given directions in advance about how to simulate Close Encounters of the First Kind.)

    Example 1: $H_a + H_b{-}H_c \rightarrow H_a + H_b{-}H_c$

This is the reaction of a hydrogen atom H_a with a hydrogen molecule H_b-H_c. H_a collides with the molecule. If H_a sticks to H_b, then H_c will be ejected from the other end and a new molecule, H_a–H_b will result. Sighting (CEFK), however, is a weak interaction of H_a with H_b but no sticking occurs; H_b–H_c remains intact.

Close Encounters of the Second Kind: Physical Evidence

An interaction between two atoms in which the atoms collide but do not stick together. Not a reactive collision—can be considered a "nonreactive" interaction. Physical evidence (CESK) is a weak interaction of H_a with H_b–H_c in which a collision occurs but no sticking takes place, so H_a merely rebounds. However, H_b–H_c absorbs some of the energy of the collision, so there is some physical evidence that something has happened. (Students acting out CESK should be given acting directions prior to the presentation to the class.)

Example 2: $H_a + H_b\text{–}H_c \rightarrow H_a + (H_b\text{–}H_c)^*$

*The H_b–H_c molecule looks like the same one as in *Example 1*, but now it has more energy and can break apart more easily into its atoms.

Close Encounters of the Third Kind: Contact

An interaction between two atoms in which the atoms not only collide but stick together to form a new molecule or ion—called a "reactive collision." Contact (CETK) is a sticky collision that forms a new chemical bond. (Students acting out CETK should be given acting directions prior to the class presentation.)

Example 3: $H_a + H_b\text{–}H_c \rightarrow H_a\text{–}H_b + H_c$

The new H_a–H_b molecule has formed! This is a successful reactive collision producing a new chemical bond.

Simply stated, if a collision is not sticky or reactive, it is not important chemically. Increasing the temperature of a chemical reaction increases the percent of effective, reactive, or sticky collisions in which new chemical bonds can be formed.

2. **Orbital Overlap: A Group Activity**

This activity requires two raw eggs, some vegetable oil, a hot plate, a kitchen skillet, an egg turner, and a good-natured student willing to fry two eggs "sunny-side-up." The student should be given instructions about the role to be played prior to the class period in which the activity is done. While the student is "performing," use this script:

Have you ever noticed how two eggs, sunny-side up in a skillet, may join together through their egg whites resulting in one double yolk bonded by the common white part of the two eggs? Then to break the bond between the two eggs to give one to a friend, the common white part has to be cut. Discuss several possible reasons why the eggs might run together and how this relates to what might happen between two atoms that are bonded. Is this a good analogy of a covalent bond between two hydrogen atoms? (Someone once said that there is no such thing as a perfect analogy.)

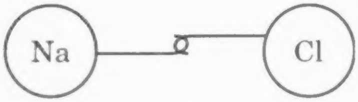
This "sunny-side-up" eggs routine is a good take-off point for discussion of orbital overlap. This seems to be a good analogy for orbital overlap in terms of everyday "stuff." It should lead to a discussion of what overlap actually means—the fusion of two half-filled atomic orbitals to form a chemical bond. Most high school chemistry texts make some attempt at this explanation; thus the activity should build a mental picture that students may use to help understand orbital overlap.

The development of bonding models is quite interesting when viewed from historical perspectives. When one thinks about how Dalton visualized chemically bonded atoms, his "fishhook" model seems very simple. In view of the fact that he knew nothing about nuclei, electrons, or valences, this model is quite logical. Arrhenius presented his theory of electrolytic dissociation in 1883 to 1887, quite some time after the introduction of Dalton's ideas of bonding.

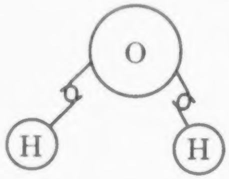
Consider:

We now know this picture of NaCl is in error since it is an ionic solid in which the bonding has significant ionic character.

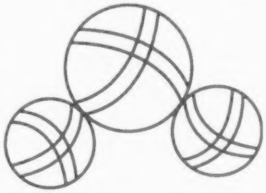
NaCl



H_2O



Is this similar to ping pong balls and Velcro?

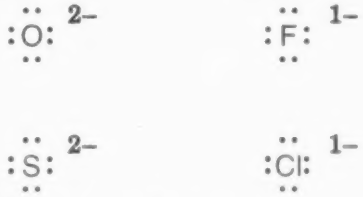


What picture does the Lewis "electron-pair" bond form? After all, we name a method of showing how the valence electrons can be arranged to explain bonding within molecules and ions. In the Lewis structure the arrangement of valence electrons can be seen around atoms in molecules and around ions in ionic compounds. Using dots to represent valence electrons, dot formulas for ions can be drawn. Remember that metals tend to lose valence electrons, acquiring a 1+ charge for each electron lost. In many cases, the ions acquire a noble gas electron configuration, a configuration recognized as highly stable. This gives rise to formulas for some common metal ions of the representative elements.

Lost 1 e⁻	**Lost 2 e⁻**	**Lost 3 e⁻**
Na^+	Mg^{2+}	Al^{3+}
K^+	Ca^{2+}	Ga^{3+}

In these examples, note that each formula implies an outer population of eight electrons—the original valence electrons have been lost, "uncovering" a stable octet of electrons in the next inner shell. Such ions are generally highly stable; they each have a noble gas electron configuration. Remember that nonmetals tend to gain valence electrons, gaining a 1– charge for each electron gained. When forming anions, nonmetals tend to acquire an octet of electrons (duet in the case of hydrogen). As in the cases of metal ions, monatomic nonmetal ions thus acquire a noble gas electron configuration. Lewis-dot formulas shown below are examples of this.

Gained 2 e⁻ **Gained 1 e⁻**



3. **Covalent and Polar Covalent Bonding**: Two children enter an ice cream store and pool their money to buy an ice cream cone to share. When the cone is in the possession of the first child, the second remains in the immediate vicinity and will not wander very far. This is like covalent bonding with equal sharing. If, on the other hand, one child is much larger than the other, the cone will likely be hogged by the larger one. This is like polar covalent bonding.

TIPS FOR THE TEACHER

Language of Chemistry

Here is a minimal glossary of terms useful in discussing chemical bonding.

coordination number number of ions of opposite charge that surround each ion in a crystal; number of atoms surrounding a given atom in a crystal.

covalent bond simultaneous attraction by two atomic nuclei for one or more electrons (usually an electron pair).

dipole interaction weak intermolecular force resulting from the attraction of oppositely charged regions of polar molecules.

double or triple covalent bond covalent bond in which two and three pairs of electrons, respectively, are shared by two atoms.

hydrogen bond relatively strong intermolecular force in which a hydrogen atom covalently bonded to a very electronegative atom is also weakly bonded to an unshared electron pair of a strongly electronegative atom in the same or adjacent molecule.

ionic bond electrostatic attraction that binds oppositely charged ions together.

Lewis-dot structure notation that depicts valence electrons as dots around the atomic symbol of the element; the symbol represents the inner electrons and atomic nucleus; also called an *electron-dot structure.*

London forces weakest type of intermolecular attraction; can be attributed to motion of electrons leading to temporary dipoles. These temporary dipoles, in turn, induce opposite temporary dipoles in adjacent molecules, thus causing weak electrostatic attraction.

metallic bond force of attraction that holds metals together; it consists of the attraction of free-moving valence electrons for positively charged metal kernels.

molecular solid solid consisting of molecules weakly attracted to each other by van der Waals forces (London forces or dipole-dipole forces) or hydrogen bonding.

network solid solid in which all atoms are covalently bonded to each other in one, two, or three dimensions.

nonpolar covalent bond homonuclear bond formed when atoms in the molecule are alike and bonding electrons are equally shared.

Octet rule atoms react by gaining or losing electrons so as to acquire the stable electron structure of a noble gas, usually eight valence electrons.

polar bond bond formed when two different atoms are joined by a covalent bond and the bonding electrons are unequally shared.

van der Waals force term used to describe the weakest intermolecular attractions; these include London forces and dipole interactions.

valence electron electron in the highest occupied energy level of an atom.

Pattern Recognition

Rules of Thumb

Rules of thumb (ROTs) are useful devices to help students remember certain facts and to increase their ability to apply general principles. Generalizations may be made and expressed as ROTs. As generalizations, ROTs have exceptions but provide a core upon which students can build, through experience. These ROTs are designed to help students recognize situations within which they can work successfully.

1. Representative element metals and nonmetals generally tend to bond ionically, forming ionic solids.
2. Nonmetals bonding with nonmetals usually bond covalently, forming molecular solids, molecular liquids, and, in some cases, network solids.
3. When an electronegativity difference is present between atoms forming a covalent bond, the resulting bond is polar.
4. An electronegativity difference between bonded atoms producing more than 50% ionic character is the arbitrary demarcation between ionic and covalent bonds.

(Rules of thumb were suggested by Bill Smith, chemistry and physics teacher at Amarillo High School, Amarillo, Texas.)

Common Student Misconceptions

1. **"Bonding must be either ionic or covalent."**

   It is a common student misconception that a bond between two atoms, A-B, is either purely covalent or purely ionic. *No compound is 100% ionic.* If the bond involves the same atoms (a homonuclear bond, A-A) then the bond must be 100% covalent because neither atom has the ability to attract the electron pair more strongly than the other. However, if the bond involves different atoms (a heteronuclear bond, A-B) the bond will have mixed covalent and ionic character. This means there will be a percent ionic character. Thus, except when the two atoms that are bonded are the same element (for example, two oxygen atoms), a bond is always partially covalent, partially ionic. The reason for this is that an electron is never completely transferred from one atom to another. The electron is shared rather than completely transferred. The sharing is a matter of degree—the concept of a polar bond. The best way to teach bonding is to show that there is a gradual progression from 100% pure covalent bond (homonuclear) to one that is about 98% ionic.

2. **"Bond energies can be reliably used to calculate heats of reaction."**

   Normally, ΔX means $X_{\text{final state}} - X_{\text{initial state}}$, where X is any common thermodynamic state function such as P, T, V, H, E, and S. However, when bond energies are calculated, ΔH means total $\text{BE}_{\text{all reactants}}$ − total $\text{BE}_{\text{all products}}$, where BE corresponds to the bond energies obtained from a table of bond energies for diatomic molecules. The ΔH obtained in this way is the energy required to break all bonds of reactant molecules minus the energy evolved when new bonds are made within product molecules.

   This value may differ from ΔH determined from heats of formation since the bonding energy tables are derived from averages of bond energies. It is therefore best to avoid using bond energies to calculate ΔH.

3. **"Mathematical or mental constructs such as electron clouds represent something solid."**

   When solid models are used to illustrate atomic scale events, it is possible to produce this misconception among students. It is quite natural for students to develop this picture since both solid models and pictures resemble solids. Caution students about taking an overly literal view of such models.

4. **"Intermolecular bonds are the same as intramolecular bonds."**

   There is quite a difference between these two types of bonds, despite their "sound-alike" nature. For example, liquid nitrogen has a low boiling point (–147 °C) due to relatively weak *inter*molecular van der Waals forces. Yet temperatures of several thousand degrees Celsius do not cause appreciable decomposition of N_2 molecules because the nitrogen atoms in the N_2 molecule are bonded with a triple covalent *intra*molecular bond. This is a strong, primary chemical bond whereas van der Waals forces are quite weak *inter*molecular forces that require relatively little energy to overcome.

5. **"Covalent bonds must be weak because covalent compounds are generally soft with low melting points (< 300 °C)."**

   Actually, this is a case of confusing *inter*molecular with *intra*molecular bonding. See the discussion above for Item 4.

Problem Solving

Number of Bonds: The Pauling-Giddings Rule—An Easy, Fast Way to Determine Lewis-Dot Structures

Determining electron dot structures for molecules is relatively easy if one knows how many bonds to draw. A powerful little bonding rule is:

$$\text{Number of bonds} = \frac{\text{electron demand} - \text{electron supply}}{2}$$

$$\text{or, } B = \frac{D - S}{2}$$

The following examples show how easy this is to use.

Example 1: How many bonds are present in a N_2 molecule?

Step 1: Each N atom "demands" an octet (8) of electrons, so $D = 2 \times 8 = 16$.

Step 2: Each N atom "supplies" the five valence electrons it has, so $S = 2 \times 5 = 10$.

Step 3: $B = (D - S)/2 = (16 - 10)/2 = 3$; there are three bonds—a triple bond. Draw the triple bond between the N atoms, then complete the octet around each N atom until all 10 "supplied electrons" are used.

Example 2: How many bonds are present in a carbonate ion, CO_3^{2-}?

Step 1: The C atom "demands" an octet (8) of electrons and each O atom "demands" an octet, so $D = (1 \times 8) + (3 \times 8) = 32$.

Step 2: The C atom "supplies" the four valence electrons it has, and each O atom "supplies" six valence electrons. However, we also must consider the net ionic charge (see next step).

Step 3: Electrons have been supplied due to the 2– charge on the ion, so
$S = (1 \times 4) + (3 \times 6) + 2 = 24$.

Step 4: $B = (D - S)/2 = (32 - 24)/2 = 4$; there are four bonds between the three oxygen and one carbon atom (from an electron-dot viewpoint, one can be regarded as a double bond; two as single bonds).

Step 5: Draw the double bond between C and one of the O atoms, then two single bonds between the C and the other two O atoms. Finally, complete the octet around each atom until all 24 "supplied electrons" are used. *(NOTE: There are three different ways to draw the C = O bond, corresponding to the three resonance structures for this ion.)*

Bonding "Coverage"

What is the minimum that should be included when teaching bonding in an introductory chemistry course? One list of minimum bonding concepts might look like this:

1. Covalent bonds and covalent network solids
2. Ionic bonds
3. Using electronegativity
4. Metallic bonds
5. van der Waals forces
6. Polar covalency
7. Lewis-dot formulas for molecules and ions
8. Valence electrons and the Periodic Table

Other bonding topics can be regarded as optional even though they may appear in the textbook. For example, concepts such as hybrid bonds are not absolutely essential.

HISTORY: ON THE HUMAN SIDE

John Dalton (1766-1844). Dalton's early career was the study of mathematics and Newtonian theories of gravity and forces. In addition, he compiled meteorological records and constructed barometers, thermometers, rain gauges, and hygrometers. His research on gases led to the formulation that in a mixture of gases each gas acts as an independent entity (law of partial pressures). Dalton was the first scientist to explain that air is a mixture of gases and not a vast "chemical solvent," as many chemists had been theorizing based on their observations of evaporation and boiling of substances. Between 1805 and 1807 Dalton began shifting his interests increasingly towards chemistry. During a series of lectures in Edinburgh, he described a new system of examining chemical combinations. In 1807 he described atoms as the indivisible particles of matter that bonded together by hooking to form all the compounds present in nature. He envisioned this hooking to be similar to fish hooks that would latch onto one another. For this and other theories, Dalton was later elected to the Royal Society. He was also instrumental in the founding of the British Association for the Advancement of Science.

Edward Frankland (1825-1899). Frankland developed a theory of valence following research into the properties of zinc methyl, a new reactive organometallic compound, and a series of alkyl-conjugated metals that had different combining powers than the metals alone. On May 10, 1852, Frankland published an article describing the fixed combining powers of elements that he termed 'atomicity.' This term was renamed *valence*. Frankland's work on valence, along with an international agreement on atomic weight values signed shortly thereafter united various theoretical schools of chemists into the field of structural chemistry.

Archibald Scott Couper (1831-1892). In 1856, Couper began research on the bromination of benzene at the laboratory of Charles Wurtz in Paris. He asked Wurtz to present a paper that he had written titled "On a New Chemical Theory" to the French Academy of Sciences. However, before his paper was presented a similar paper was presented by Friedrich Kekulé. By the time Couper's paper was read, Kekulé had claimed priority for the formulas for bromobenzene (C_6H_5Br) and p-dibromobenzene ($C_6H_4Br_2$). Couper, after complaining to Wurtz about the delay in presenting his paper, was dismissed from the laboratory and returned to his native Scotland. His subsequent research on carbon showed that it has a valence, or combining power, of two or four, forms valence links with other carbon atoms, and thereby forms carbon chains. Couper represented these links with lines and in 1858 introduced the first "ring formula" to organic chemistry.

Linus Pauling (1902-). Pauling began his studies as an undergraduate at Oregon State Agricultural College. In order to pay for his classes, he taught quantitative analysis and offered tutorials on freshman chemistry. By 1920, when he was a senior, Pauling was offering a seminar on the electronic nature of the chemical bond. He based this course on Lewis and Langmuir's theories of chemical bonding published in 1916. Pauling attended the California Institute of Technology beginning in 1922. At the time Caltech was one of the first institutions using the new methods of X-ray crystallography to analyze structures, and while still a graduate student, Pauling published seven papers on crystal structures. By 1928, after studying in Europe for a year, Pauling was doing frontier research on chemical bonding. Beginning in 1931 he published a seven-part series titled "The Nature of the Chemical Bond." New rules for determining bond lengths, the angles between bonds, magnetic moments and other molecular properties were described using the concept of *resonance*. Pauling's theory held that molecules can be represented by a linear combination of wave functions. He thus transformed the field of chemistry by applying quantum theory and quantum mechanics to chemical structure and bonding. Pauling has continued to study molecular structure with emphasis on the structure of protein molecules from living tissue. Through his success in determining protein structures, Pauling established himself as a founder of modern molecular biology.

In 1954, Pauling received the Nobel Prize for his research into the nature of the chemical bond. In addition to his scientific work he has been an outspoken advocate of peaceful resolution of conflict. Such activities resulted in his receipt of a second Nobel Prize, the Peace Prize, in 1963.

G. N. Lewis made significant contributions to ideas about the chemical bond. A brief biography is in Masterton, W. L., and Hurley, C. N. (1989). *Chemistry: Principles and reactions* (pp. 272-273). New York, NY: Saunders.

HUMOR: ON THE FUN SIDE

1. Capitalism reigns in the atomic world. The unequal sharing of electrons is the central concept of bonding. One of history's most sage remarks concerning sharing was uttered by Winston A. Churchill, Prime Minister of England during World War II, which was also once used in a beer advertisement:

   > *The inherent virtue of socialism*
   > *is the equal sharing of miseries;*
   > *the inherent vice of capitalism*
   > *is the unequal sharing of Budweiser.*

   What did Churchill actually say? (Yes, the unequal sharing of wealth!) This was, indeed, a very wise remark on Churchill's part. Some people do have more Budweiser, wealth, luck, and wisdom than do other people. Some atoms, on the other hand, are more electron-wealthy than others—they are said to be more electronegative! Fluorine is the most wealthy, the most capitalistic of all atoms. Electrons are not shared equally between atoms unless those atoms are the same, so unequal sharing is the normal situation. Capitalism reigns supreme in the atomic world!

2. Elements 4 and 18 (from the alkaline earth and noble gas families) form a 1:1 compound that is polar. What is the compound's color and formula? *Answer: BeAr. Since it is polar, it must be white.*

3. **Question**: What is the name of 007's Eskimo cousin? *Answer: Polar Bond. (CHEM 13 NEWS, April 1975, p. 894)*

4. **Limericks**:

   a. Bond formation can me unwell
   (Electrons join up in a shell)
   For it doesn't seem fine
   To say they combine
   When we know very well they repel!
   (CHEM 13 NEWS, May 1988, p. 7)

   b. While he was out walking alone,
   Old Edward decided to clone.
   He found himself doubled,
   But not the least troubled,
   For two Ed's are better than one!

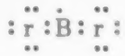
   c. There's a law that just simply can't be
   About time and space fantasy.
   But Albert Einstein
   Theorized it just fine,
   Because E over M is squared C.
   (CHEM 13 NEWS, May 1981, p. 9)

5. **Rhyme**:

   Chemistry is full of rules
   Atoms, moles, and molecules.
   Some of which might fit together,
   Maybe now, and likely never.
   (CHEM 13 NEWS, January 1972, p. 380)

6. **Message on a T-shirt**: Together we could make beautiful molecules.

7. The adjacent Lewis-dot diagram for Br_2 was one of several similar variants submitted on a recent test by students who shall remain anonymous.
   (CHEM 13 NEWS, December 1991, p. 2)



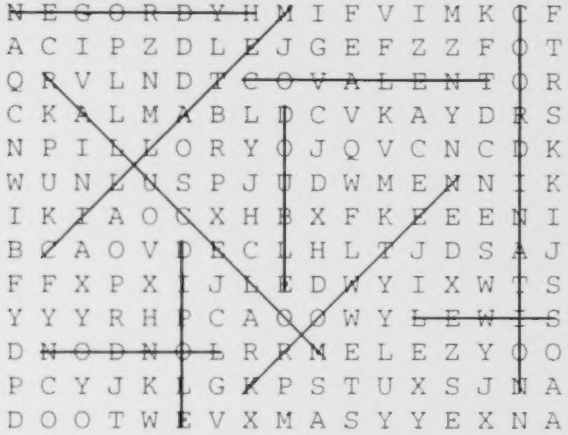
:r:B:r:

8. **CHEMTOON**: H :F: We seem to be poles apart.
   (CHEM 13 NEWS, September 1983, p. 15)

9. **ON HYBRIDS**

   Beryllium gave many a plea
   'Cause it's valence electrons weren't free
   And it couldn't combine
   till a thought came to mind
   And it went and invented sp!
   Boron's not satisfied yet
   Its demands really cannot be met
   Though it tries hard to be
   Like He or Ne
   sp^2 is the best it can get.
   Carbon's like one seen before
   His shells couldn't hold any more
   Till his friends lent a hand
   showed him how to expand
   And sp^3 lets him join up with four!
   (CHEM 13 NEWS, September 1981, p. 3)

10. Word Search (see *Appendix* for master copy)



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N E G O R D Y H M I F V I M K C F
A C I P Z D L E J G E F Z Z F O T
Q R V L N D T C O V A L E N T O R
C K A L M A B L D C V K A Y D R S
N P I L L O R Y O J Q V C N C D K
W U N L U S P J U D W M E N N I K
I K I A O C X H B X F K E E E N I
B C A O V D E C L H L T J D S A J
F F X P X I J L E D W Y I X W T S
Y Y Y R H P C A O O W Y L E W I S
D N O D N O L R R M E L E Z Y O O
P C Y J K L G K P S T U X S J N A
D O O T W E V X M A S Y Y E X N A
```

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Type of bond that holds metals together.
2. _______ bond, simultaneous attraction by two atomic nuclei for one or more electrons.
3. Weak intermolecular force between polar molecules.
4. Type of solid in which all atoms are covalently bonded.
5. Type of solid in which molecules are weakly attracted to each other.
6. Relatively strong type of bond between hydrogen and a highly electronegative atom such as O or N.
7. Weakest type of intermolecular attraction.
8. Name for electron-dot notation.
9. Type of covalent bond in which two pairs of electrons are shared.
10. Number denoting number of ions of opposite charge surrounding an ion in a crystal.

Answers: 1. METALLIC 2. COVALENT 3. DIPOLE 4. NETWORK 5. MOLECULAR 6. HYDROGEN 7. LONDON 8. LEWIS 9. DOUBLE 10. COORDINATION

11. See cartoons at end of module.

MEDIA

1. *Solo-Learn* series of auto-tutorial programs consisting of a filmstrip, a cassette commentary, student worksheets and teacher's guide are available in VHS or 16-mm film format for purchase or rental from Ward's Natural Science Establishment, Inc., 5100 West Henrietta Road, P.O. Box 92912, Rochester, NY 14692-9012; (800) 2660.

   a. *Introduction to Chemical Bonding* presents the basics of chemical bonding including valence electrons, Lewis dot structures, electronegativity, ionic, covalent, and polar covalent bonding.

   b. *Polar Covalence* presents the details of polar covalent bonds including properties associated with bond types, polar molecule behavior in electric fields, predicting degree of polarity, orbital overlap with regard to bond formation and strength, and prediction of properties due to bond polarity.

c. *Bond Type and Properties of Matter* describes bonding in terms of electrostatic attraction and repulsion, and correlates properties of substances with the bonding and structure of elements and compounds.

These autotutorial programs are thorough in coverage and particularly useful for individual or small group work. They may be used for presentation to large groups also.

2. CHEM Study Films and Videos (1/2 -inch VHS) All the original CHEM Study films have been revised and brought up to date. They are also available in convenient video format. Highly recommended—particularly the first and third titles listed below. *Crystals and their Structures* also fits well in a Condensed States unit. Available from Ward's Natural Science Establishment, P.O. Box 92912, Rochester, NY 14692-9012, (800) 962-2660, in VHS format or 16-mm film format for purchase or rental.
   a. *Chemical Bonding* explains chemical bonding in terms of the electric interactions that cause bonding in the hydrogen molecule. 16 min. Recommended for average or better students.
   b. *Crystals and their Structures* introduces the properties of crystals, their well-defined diffraction patterns in interaction with X-rays, and their description as regular, repeating arrangements of atoms. 22 min. Recommended for above average students.
   c. *Electric Interactions in Chemistry* develops Coulomb's law experimentally in a way that average or better students can grasp. The experiment used is somewhat unique. 21 min. Recommended for average or better students.
   d. *Shapes and Polarities of Molecules* develops the concept of molecular polarity. Polar molecules are used to explain differences in solubility, electrical conductivity, and chemical reactivity. 18 min. Recommended for average and above average students.
3. *Chemical Bonding and Atomic Structure*, 23 min. video available from Coronet/NTI Film and Video, 108 Wilmot Road, Deerfield, IL 60015; (800) 621-2131; (708) 940-3640 (FAX).
4. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).
   a. *MolVib: Visualizing Molecular Vibrations*, by Daniel Huber. Vol. III C, No. 1, for the Apple Macintosh.
   b. *Evolution of Bonding Theory*, by David Whisnant. Vol. II A, No. 1, for the Apple II computer.
5. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   For the Apple II computer running on ProDOS: AR 303
6. Videodisc published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, Wl 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Experimenting with Electrical Charges," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

Links/Connections

WITHIN CHEMISTRY

Bonding may be considered the central concept in chemistry since it determines the microscopic structure of all matter. The structure of matter, in turn, determines the properties it possesses. The linkages between chemical bonding and atomic structure, periodicity, reaction kinetics, chemical equilibrium, and chemical thermodynamics are especially strong. Bonding in polymers – textiles, plastics, *etc.* – is economically important. Disruption of the bonding in ozone by the action of free radicals has led to depletion of ozone in the upper atmosphere.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Since bonding determines the properties of substances—both macroscopic and microscopic—it is definitely linked to other scientific disciplines. This linkage is more obvious in some cases than in others. For example, microbiology, medicine, chemical engineering, pharmacology and toxicology are closely linked to chemical bonding. Much of our understanding of food preservation and the use of food additives (both helpful and harmful) is based on chemical bonding. Hydrogen bonding is important for the functioning of DNA

TO THE CONTEMPORARY WORLD

Personal

Most people are familiar with "bonding" involving magnetism, Velcro, familial bonding, *etc.* Chemical bonding is less conspicuous, but more pervasive. In fact, *everything* is chemically bonded including people and other organisms. Personal linkage to chemical bonding is so widespread that no one consciously considers it or wonders about it. Introducing some "wondering why" about the role of chemical bonding in daily life is highly appropriate and necessary.

Community

Knowledgeable Individuals

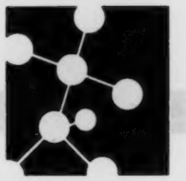
University, college, and community college teachers, professional chemists, forensic chemists, chemical engineers, medical doctors, experienced chemistry teachers, *etc.*, can all provide assistance and perspectives. Most local sections of the American Chemical Society have an education committee that can also help.

Other Community Resources

Public libraries, community utility departments, hospitals, and police departments may be useful resources.

Societal

Local newspapers are sources of chemical based problems that are useful to produce student awareness.



Extensions

1. Research the significance of hydrogen bonding in the role of water in climate, living organisms, and life in general.
2. Research the explanation of metallic properties, the location of metals in the Periodic Table, and the properties of alloys in terms of the metallic bond.
3. Research the explanation of the properties of covalent network solids in terms of the covalent bonding involved.
4. Research the trend in bond type from pure covalent to ionic on the basis of electronegativity and electron affinity.
5. Design and perform an experiment to investigate the relation between solubility and charge separation in both solute and solvent.
6. Design and perform an experiment to investigate the formation of a precipitate when two solutions are mixed.

References

Module developed by Royace Aikin, Daniel Kallus, and Russell D. Larsen, the Texas Team.

Boorse, H. A., and Motz, L. (Eds.). (1966). *The world of the atom, Vols. 1 & 2.* New York, NY: Basic Books, Inc.

All you ever wanted to know about the atom and more. An excellent resource for historical information and reprints of all or parts of original papers by "giants" who developed our current understanding of the structure of matter.

Campbell, J. A. (1978). *Chemistry: The unending frontier.* Santa Monica, CA: Goodyear.

Chemistry is presented in an unusual and interesting fashion. Some of the information, particularly statistical data, is dated, but the chemistry presented supports a Science-Technology-Society perspective.

Griswold, E. (1968). *Chemical bonding and structure.* New York, NY: Raytheon Education.

One of a genre of paperback books written during the 1960s and 1970s that dealt with a single major topic usually encountered in general chemistry. Clear exposition of chemical bonding and its relationship to structure.

Pimentel, G. (Ed.). (1966). *Chemistry—An experimental science.* San Francisco, CA: W. H. Freeman and Co.

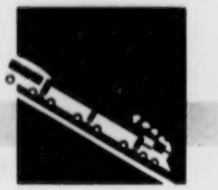
Some overhead transparencies are adapted from figures in this classic text.

Schulz, E. (Chair). (1983). *The tangent sphere model: An analog to chemical structure.* Princeton, NJ: Woodrow Wilson National Fellowship Foundation.

A valuable collection of demonstrations, laboratory experiments, and expositions. Developed by high school chemistry teachers for high school chemistry teachers.

Woodburn, J. H. (1976). *Taking things apart and putting things together.* Washington, DC: American Chemical Society.

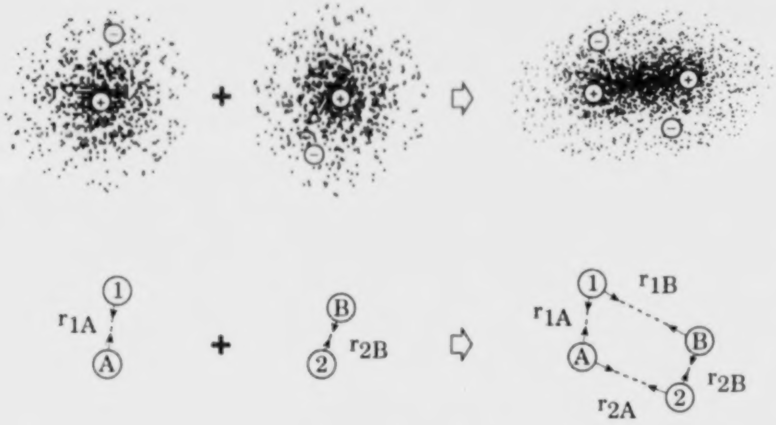
This paperback was written for the American Chemical Society's centennial anniversary. Interesting, informative reading for high school students and the general public. Highly recommended and still available.



Appendix

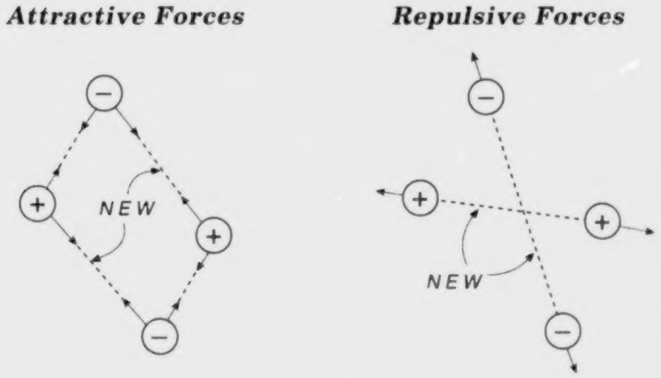
- **Transparency Masters**
  1. The Formation of a Molecule of Hydrogen, H_2 (Covalent Bonding)
  2. Attractive and Repulsive Forces When Helium Atoms Approach (Covalent Bonding)
  3. Schematic Representation of the Interaction Between Two Atoms (Covalent Bonding)
  4. Electron Distribution in Various Bond Types (Covalent ↔ Ionic Bonding)
  5. The Boiling Points of Some Hydrides (Hydrogen Bonding)
  6. Types of Network Silicates (Covalent Networks)
  7. Slippage of Planes of Metal Atoms (Metallic Bonding)
  8. Pauling Fractional Ionic Character (Covalent ↔ Ionic Bonding)
  9. Word Search
- **Humor**

The Formation of a Molecule of Hydrogen, H_2



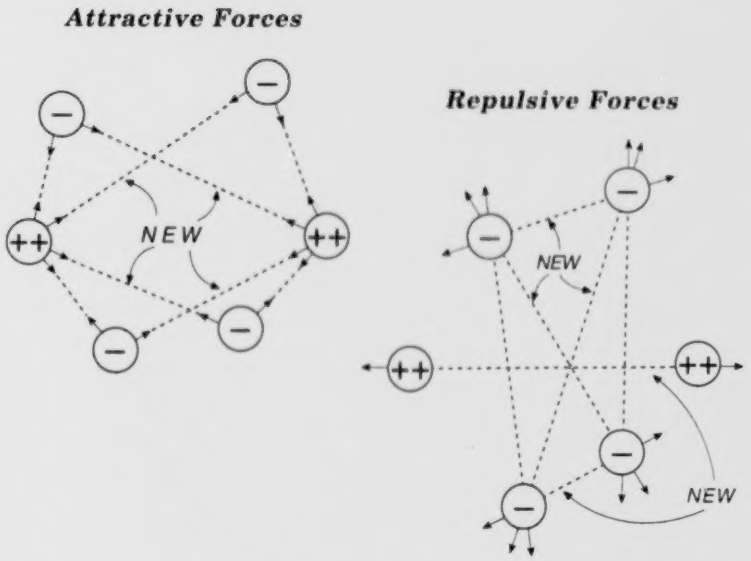
1
r_{1A}
A
+
B
2
r_{2B}
1
r_{1B}
r_{1A}
A
B
r_{2B}
r_{2A}
2

Attractive and Repulsive Forces in the Hydrogen Molecule


Attractive Forces

Repulsive Forces

Attractive and Repulsive Forces When Helium Atoms Approach

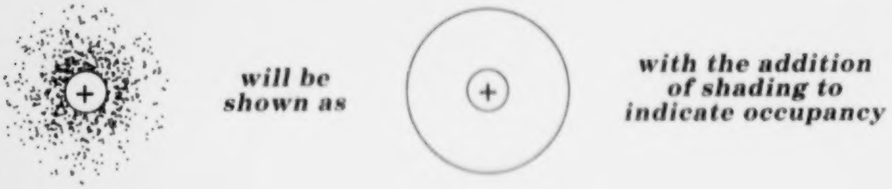


Attractive Forces

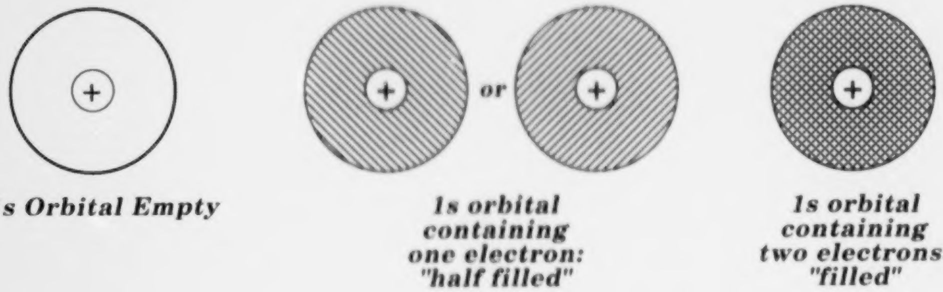
Repulsive Forces

Schematic Representation of the Interaction Between Two Atoms

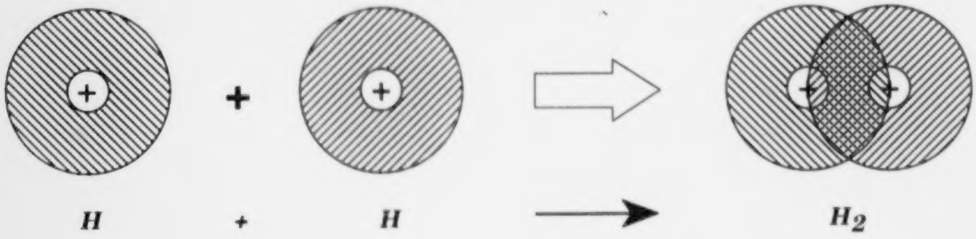
A. A Simplified Representation of a 1s Orbital:


+
will be shown as
+
with the addition of shading to indicate occupancy

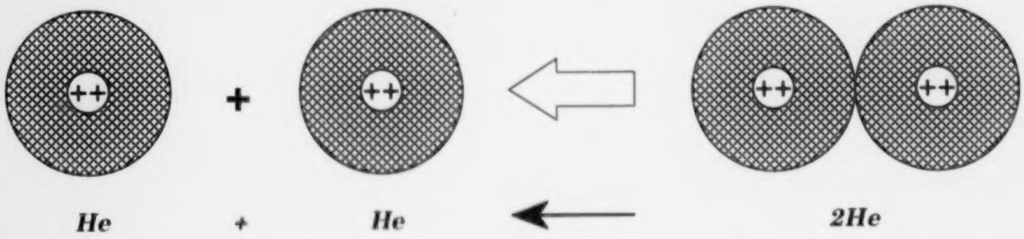
B. A Simplified Representation of Orbital Occupancy:


+
1s Orbital Empty
+
or
+
1s orbital containing one electron: "half filled"
+
1s orbital containing two electrons: "filled"

C. Overlap and Bonding of the Hydrogen Molecule:

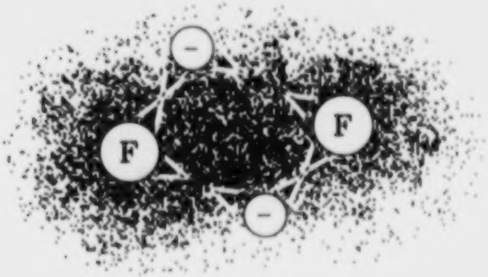


D. Absence of Overlap for Two Helium Atoms:

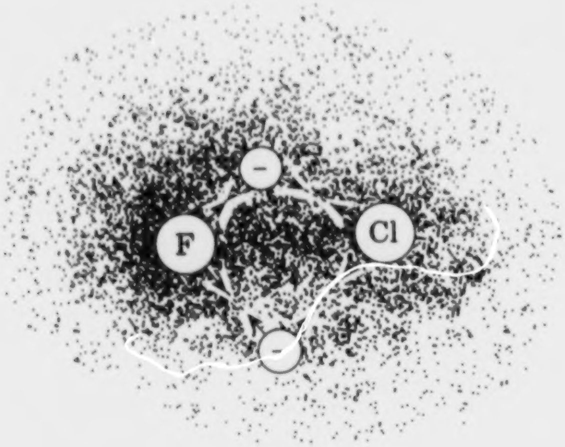


Electron Distributions in Various Bond Types

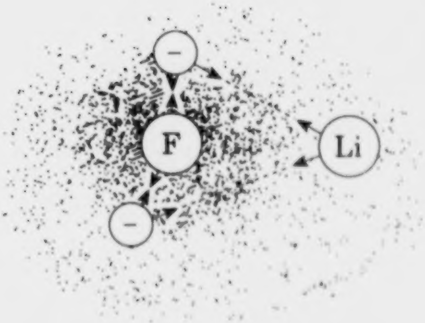
Covalent Bond
F – F



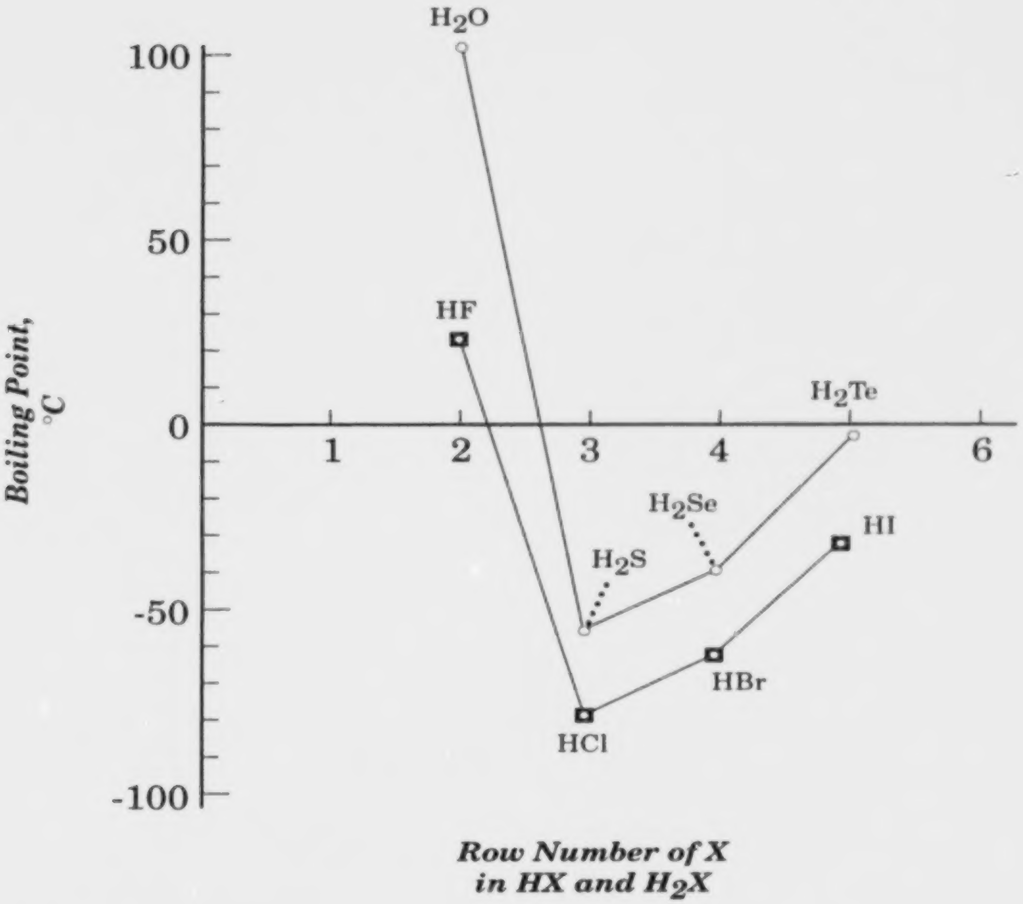
Covalent Bond with
Partial Ionic Character
F – Cl



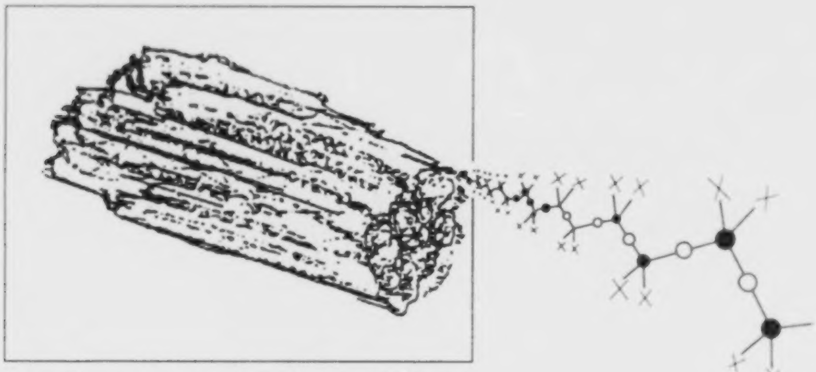
Ionic Bond
F – Li



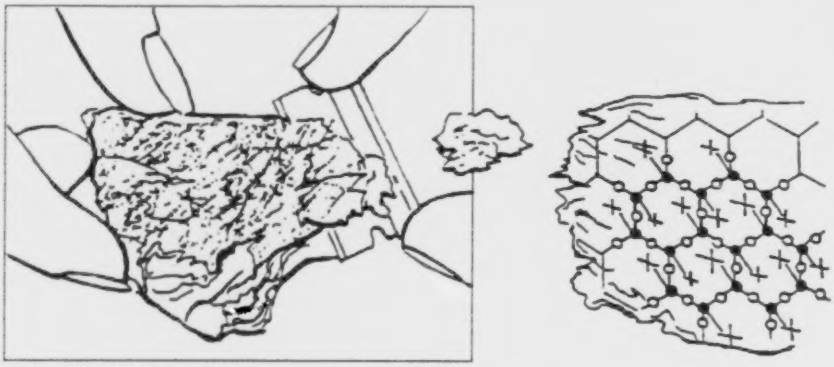
The Boiling Points of Some Hydrides



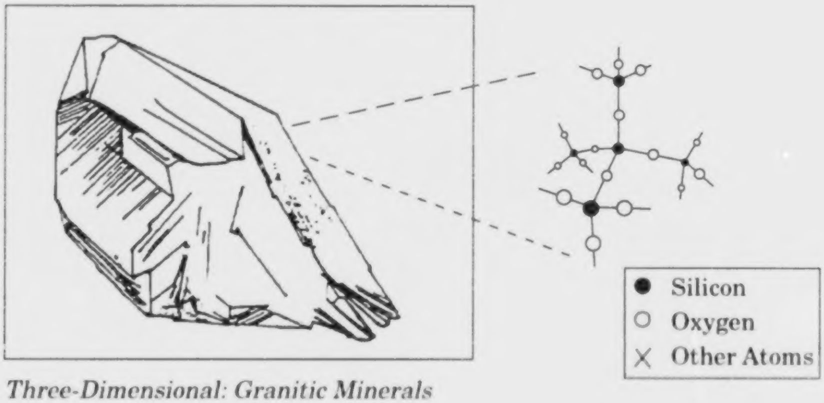
Types of Network Silicates



One-Dimensional: The Asbestos Minerals

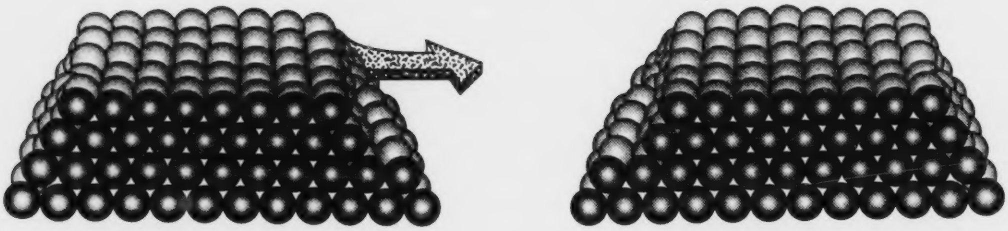


Two-Dimensional: The Mica and Clay Minerals

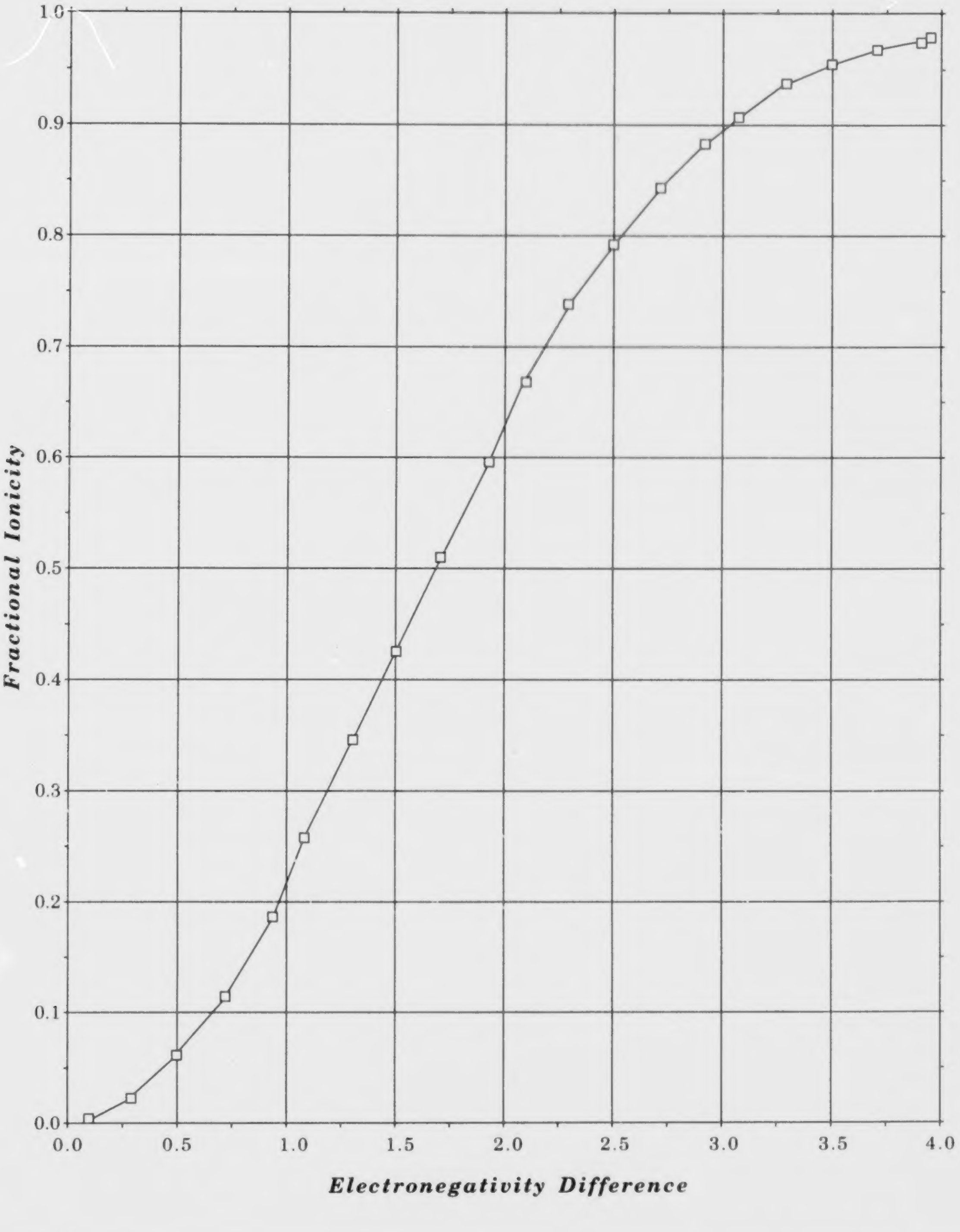

● Silicon
○ Oxygen
× Other Atoms

Three-Dimensional: Granitic Minerals

Slippage of Planes of Metal Atoms



Pauling Fractional Ionic Character



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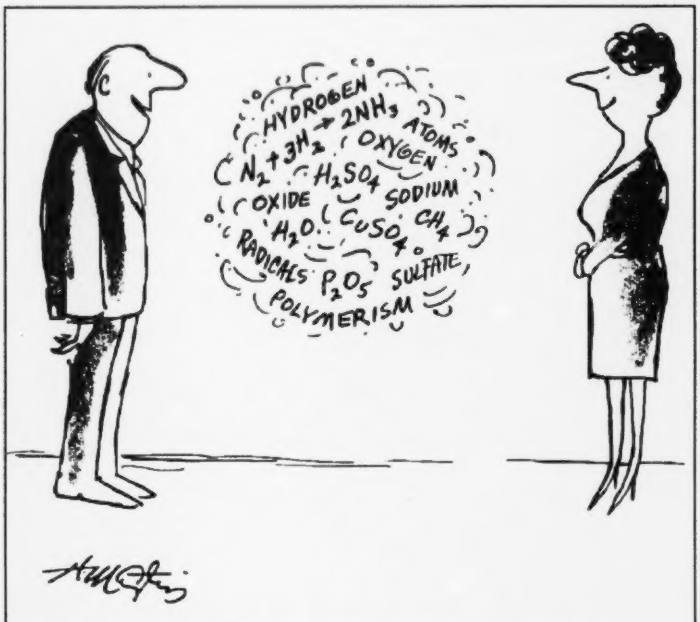
Word Search

N E G O R D Y H M I F V I M K C F
A C I P Z D L E J G E F Z Z F O T
Q R V L N D T C O V A L E N T O R
C K A L M A B L D C V K A Y D R S
N P I L L O R Y O J Q V C N C D K
W U N L U S P J U D W M E N N I K
I K I A O C X H B X F K E E E N I
B C A O V D E C L H L T J D S A J
F F X P X I J L E D W Y I X W T S
Y Y Y R H P C A O O W Y L E W I S
D N O D N O L R R M E L E Z Y O O
P C Y J K L G K P S T U X S J N A
D O O T W E V X M A S Y Y E X N A

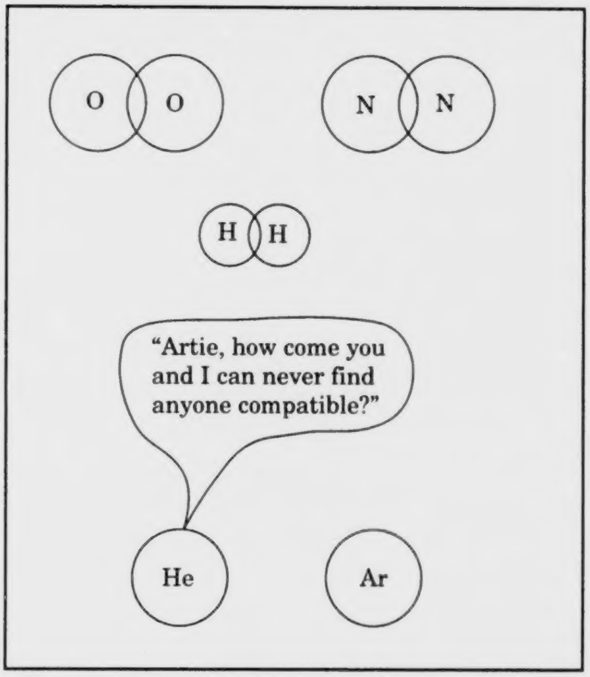
Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Type of bond that holds metals together.
2. _______ bond, simultaneous attraction by two atomic nuclei for one or more electrons.
3. Weak intermolecular force between polar molecules.
4. Type of solid in which all atoms are covalently bonded.
5. Type of solid in which molecules are weakly attracted to each other.
6. Relatively strong type of bond between hydrogen and a highly electronegative atom such as O or N.
7. Weakest type of intermolecular attraction.
8. Name for electron-dot notation.
9. Type of covalent bond in which two pairs of electrons are shared.
10. Number denoting number of ions of opposite charge surrounding an ion in a crystal.


HYDROGEN
$N_2 + 3H_2 \rightarrow 2NH_3$
ATOMS
OXYGEN
H_2SO_4
OXIDE
SODIUM
H_2O
$CuSO_4$
CH_4
RADICALS
P_2O_5
SULFATE
POLYMERISM

"As time goes by, Claralou, I think the chemistry between us is getting better."

By permission of Henry R. Martin.


O
O
N
N
H
H
"Artie, how come you and I can never find anyone compatible?"
He
Ar

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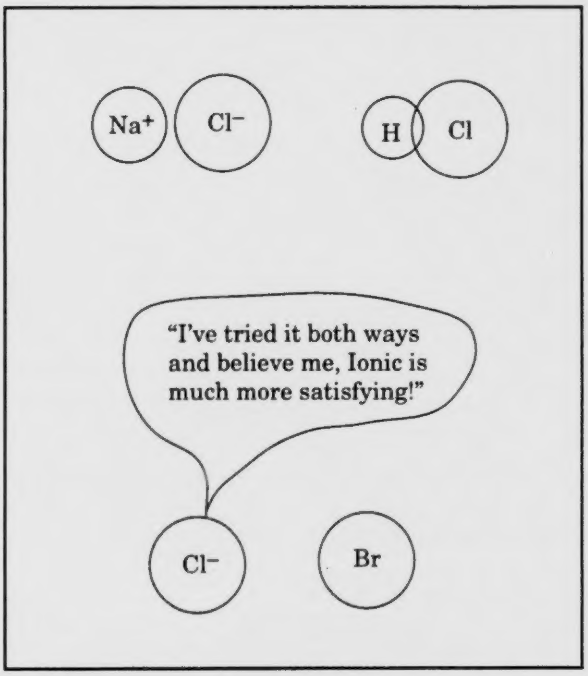
Na+
Cl−
H
Cl
"I've tried it both ways and believe me, Ionic is much more satisfying!"
Cl−
Br

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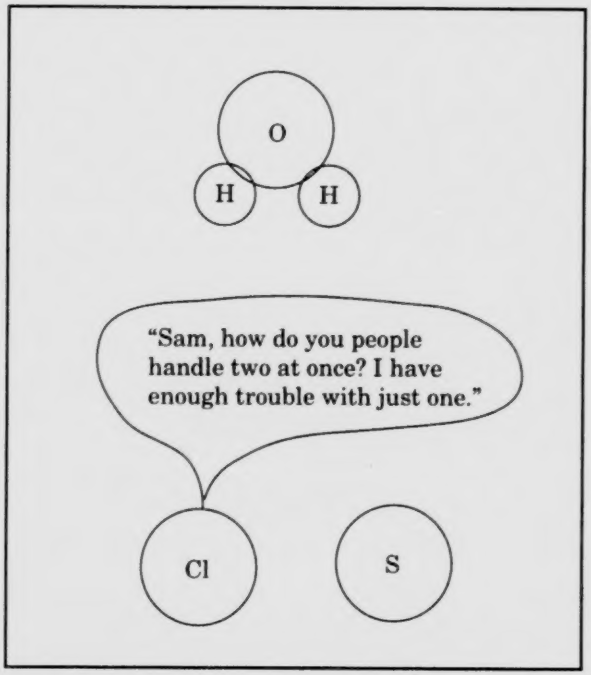
O
H
H
"Sam, how do you people handle two at once? I have enough trouble with just one."
Cl
S

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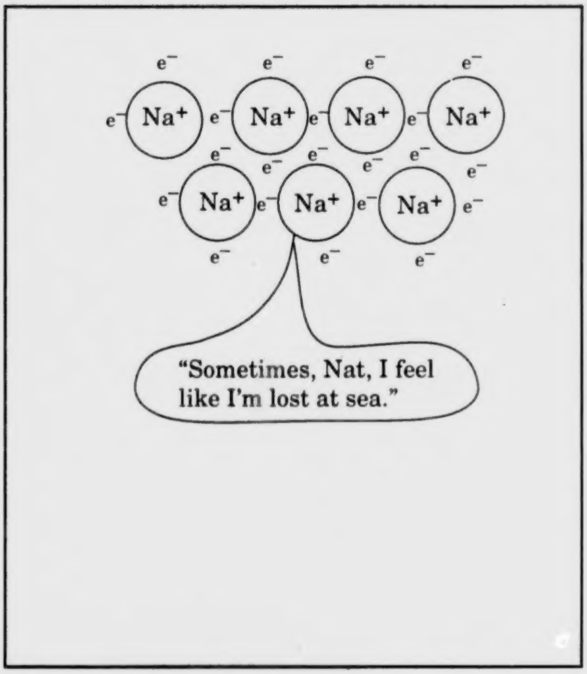
e⁻
Na⁺
"Sometimes, Nat, I feel like I'm lost at sea."

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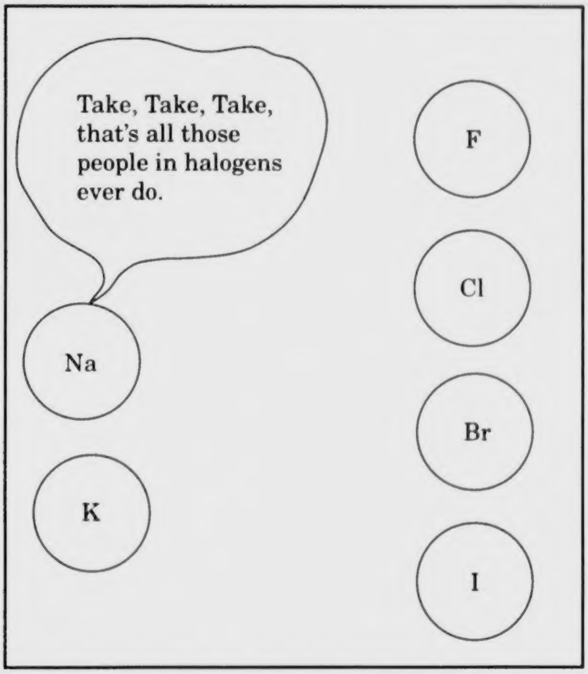
Take, Take, Take, that's all those people in halogens ever do.
F
Cl
Br
I
Na
K

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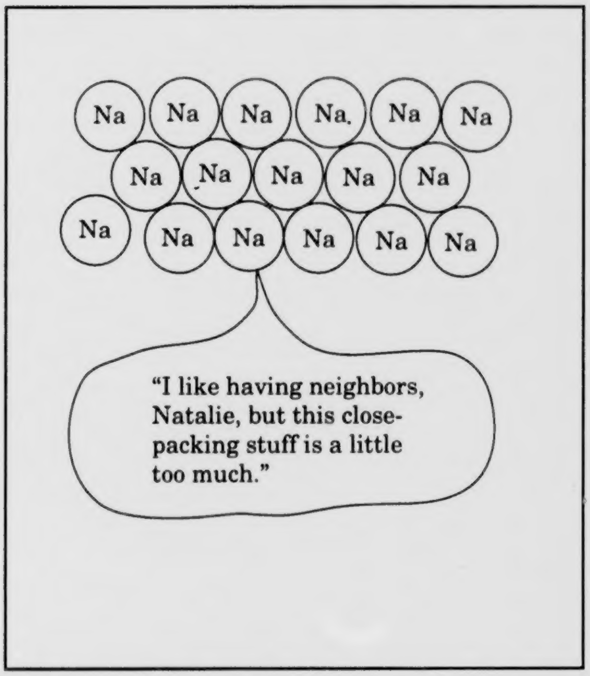
Na Na Na Na Na Na
Na Na Na Na Na
Na Na Na Na Na Na
"I like having neighbors, Natalie, but this close-packing stuff is a little too much."

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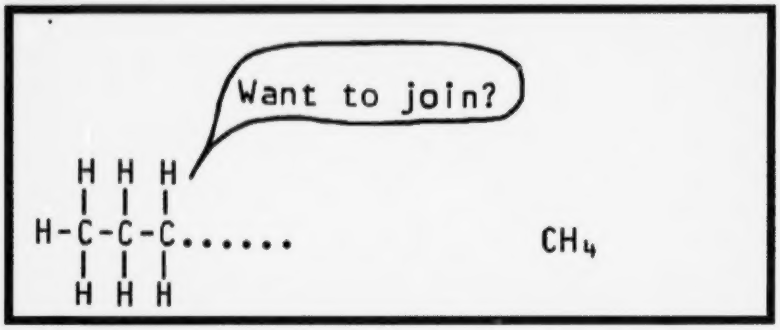
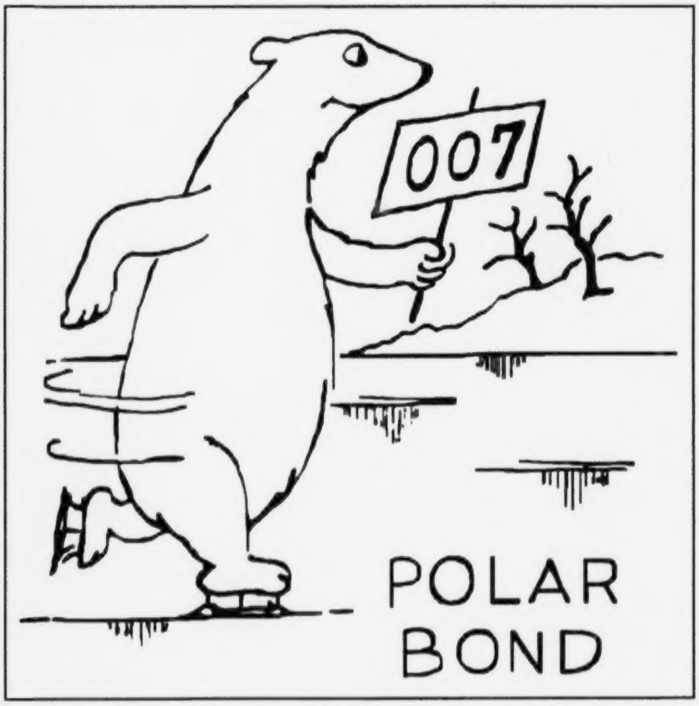

Want to join?
H H H
H-C-C-C......
H H H
CH_4

CHEM 13 NEWS, January 1984, p. 10. Reprinted with permission.


007
POLAR
BOND

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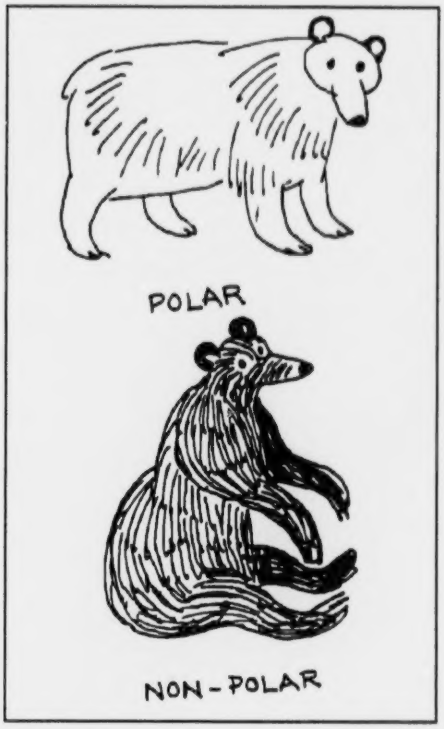
POLAR
NON-POLAR

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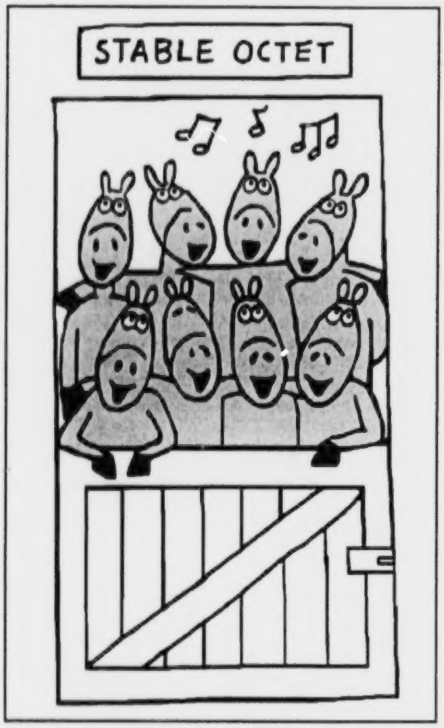
STABLE OCTET

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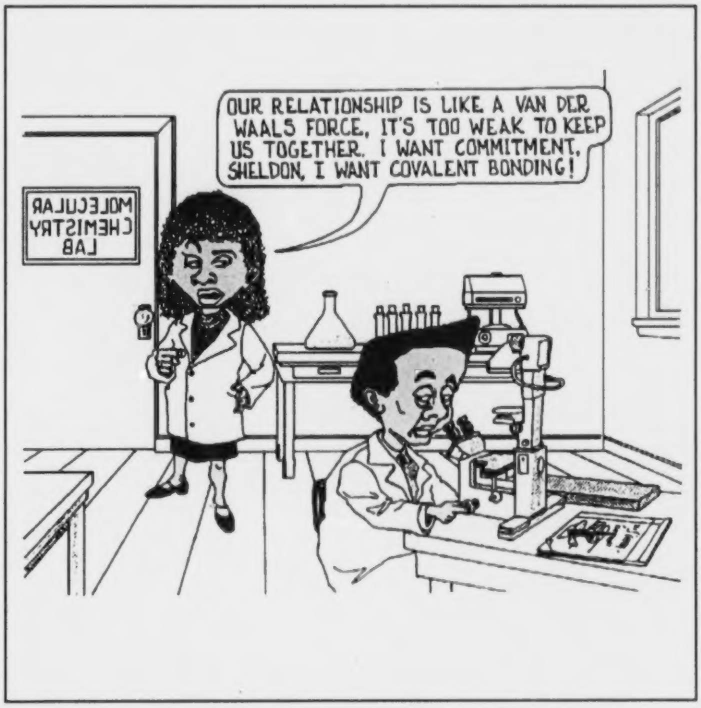
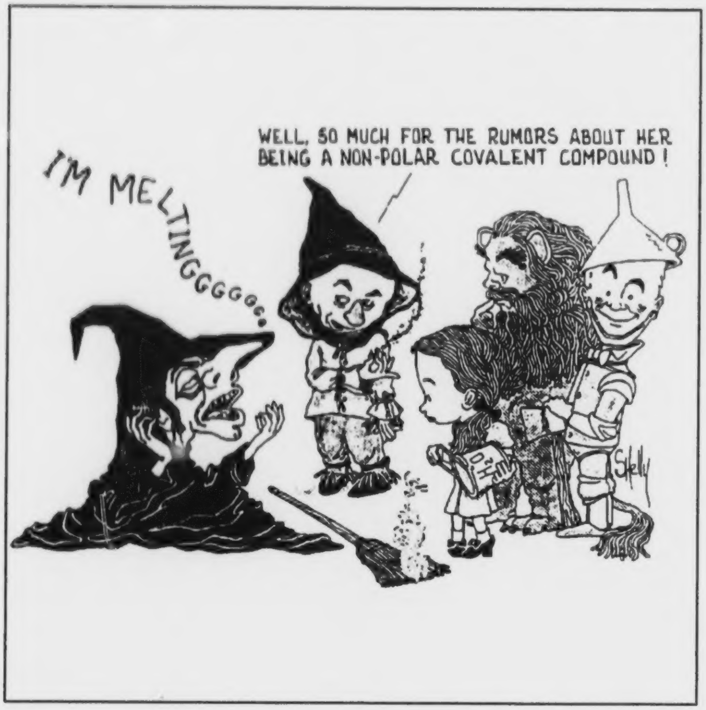

OUR RELATIONSHIP IS LIKE A VAN DER WAALS FORCE, IT'S TOO WEAK TO KEEP US TOGETHER. I WANT COMMITMENT, SHELDON, I WANT COVALENT BONDING!
MOLECULAR CHEMISTRY LAB

Cartoon by Shelly Fischman. Reprinted with permission.


I'M MELTINGGGGGG
WELL, SO MUCH FOR THE RUMORS ABOUT HER BEING A NON-POLAR COVALENT COMPOUND !
H2O
Skelly

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A SourceBook Module

Version 2.0 1994

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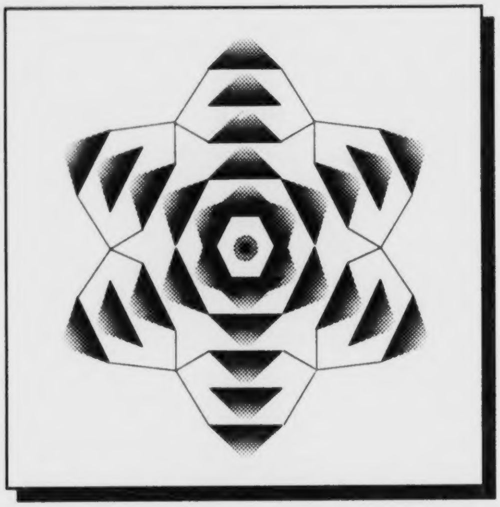
FAX: (914) 654-5387

CONDENSED STATES



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

The model used to describe the behavior of gases is called the kinetic theory. The main difference between the gaseous model and that for condensed states (liquids and solids) is that, in condensed states, there are short range forces between particles which cause them to stay in contact—to be attracted to one another. In some substances, hydrogen for example, these attractive forces are extremely weak and thus the substance turns liquid and then solid only at very low temperatures. The amount of empty space in the condensed states is small but not zero.

In liquids, particles are envisioned to be in close contact with one another, and with nearest neighbors arranged around one another in fairly regular ways. For a particle in the liquid to escape and become a gaseous particle, the addition of enough energy to overcome the attractive forces between the particles is required.

Crystalline solids differ from liquids in their degree of bonding regularity. Solids are thought to be extraordinarily regular in their arrangement such that, knowing the positions of just a few particles allows the precise prediction of the positions of nearly all others. We speak of a crystal lattice, an imaginary regular geometric array of points such that each lattice point serves as a marker for the location of a solid particle.

In order for a solid to melt and become a liquid, energy must be added to overcome the highly directional attractive forces which hold the solid particles in a regular array. The energy for melting is nearly always lower than the energy for boiling.

PLACE IN THE CURRICULUM

This unit can follow gas laws and discussions of kinetic energy. It also can precede or blend easily with thermochemistry. A discussion of the unit cell, coordination numbers, and crystal structures can fit in appropriately. References to phase changes are appropriate when studying LeChatelier's principle.

CENTRAL CONCEPTS

1. Solids and liquids are the two condensed states of matter.
2. Solids have definite shapes and volumes. Liquids have definite volumes but assume the shapes of their containers.
3. Density is an important characteristic property of liquids and solids.
4. As solids or liquids are warmed, the kinetic energy of their particles increases, increasing particle motion. As solids melt or liquids boil, the potential energy of the particles increases as the forces of attraction are disrupted (and vice versa).
5. Liquids in a closed container can achieve a condition known as dynamic equilibrium where the rate of vaporization equals the rate of condensation.
6. Solids in a closed container can achieve a condition known as dynamic equilibrium where the rate of melting equals the rate of freezing.
7. The pressure exerted by the gaseous phase of a liquid (or solid) at dynamic equilibrium is called the equilibrium vapor pressure. Vapor pressure values are influenced by the bond strengths (or attractive forces) between the liquid (or solid) particles.

8. The boiling point of a liquid is the temperature at which the vapor pressure of the liquid equals the prevailing external (atmospheric) pressure. At the melting point, the vapor pressures of the solid and liquid phases are equal.
9. The surface tension of a liquid is caused by unbalanced forces on the surface particles and makes the surface appear to be elastic. Surface tension is responsible for capillary rise.
10. The viscosity of a fluid is its resistance to flow.
11. Pure crystalline solids have well-defined melting (freezing) points which are greatly influenced by the type and strength of bonding among particles in the crystal.

RELATED CONCEPTS

1. Gases
2. Bonding
3. Atomic and ionic radii
4. Potential and kinetic energy
5. LeChatelier's principle
6. Kinetic molecular theory
7. Chemical forces (bonds) *vs.* physical attractive forces
8. Coordination numbers
9. Archimedes' Principle

RELATED SKILLS

1. Preparing a graph.
2. Dealing with three dimensional geometry.
3. Determining the slope of a line.
4. Reading balance, thermometer, and graduated cylinder.

PERFORMANCE OBJECTIVES

After completing their study of condensed states, students should be able to:

1. list the two condensed states of matter and describe characteristics of each condensed state.
2. graph mass and volume data for an unknown solid sample and determine the density of the solid from the slope of the line.
3. describe the physical changes and energy changes which take place as substances melt, boil, condense, and freeze.
4. offer an appropriate model and explanation for dynamic equilibrium in a closed container where a solid/liquid or liquid/gas system exists.
5. define melting point and boiling point in terms of vapor pressure.
6. relate the vapor pressure of a solid or liquid substance to the attractions between its particles.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: Density, A Straight Line Function

Introduction

The density of a pure substance does not depend upon the amount of that substance present. Like its temperature or melting point or color, density is an intensive property of a liquid or solid. A ton of lead has the same density as a milligram of lead (11.36 g/cm^3). Common units for density are grams per cm^3 for solids, grams per mL for liquids, and grams per L for gases. In this activity we will work with solids.

Purpose

The formula for density will be developed from experimental data by graphical interpretation. To calculate the densities of aluminum metal and iron metal by two methods and compare the methods.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Use normal safety precautions with this experiment.

Procedure

Work in pairs to accomplish the following steps:

1. You will be given 6-8 different-sized metallic objects. Tie a thread to each metallic object as indicated by your instructor. This will enable you to hang the object from the balance. Determine the mass of the metallic object as it hangs from the balance, and record the mass in the air to the nearest 0.01g.
2. To find the volume of each object, there are two possible procedures, the volume displacement method and the Archimedes' weight displacement method. If you choose or are instructed to use the volume displacement method, go to Step 3. If the Archimedes' weight displacement method is selected then go to Step 4.
3. Add enough water to the graduated cylinder so that the largest metallic object can be wholly submerged. For each metallic object take an initial water level reading to the nearest 0.2 mL, and record. Carefully lower the metallic object by its thread into the water in the graduated cylinder until it is completely immersed. Read the water level in the graduated cylinder and record to the nearest 0.2 mL. Be sure to take new initial and final readings with each metallic object. Figure 1 is a sample data table to be completed.

Sample No.	1	2	3	4	5	6	7	8
Mass of object (in air)								
Init. volume (of water)								
Final volume (of water)								
Initial-final volume								
Volume of metallic object								

Figure 1. Data table for volume displacement method.

4. The Archimedes' weight displacement method is easy to accomplish using a balance equipped with a platform (Figure 2). Add water to a 250-mL beaker and then place the beaker on the platform. Hang the object from the balance by the thread so that it is completely submerged in water. Take care not to let the metallic object touch the sides of the beaker. Read the balance to the nearest 0.01 g and record the data in your data table (Figure 2). Repeat this procedure for each metallic object.

   STUDENT NOTE: Archimedes' principle states that a body partially or totally submerged in a liquid is buoyed up by a force equal to the weight of the displaced liquid. This accounts for the apparent loss of mass. The density of water is 1.0 g/cm^3. If an object displaces 25-mL water, the apparent loss of mass is 25 g. Consequently the difference between mass in air and mass in water is equivalent to the volume of water displaced and the unknown volume of the object itself.

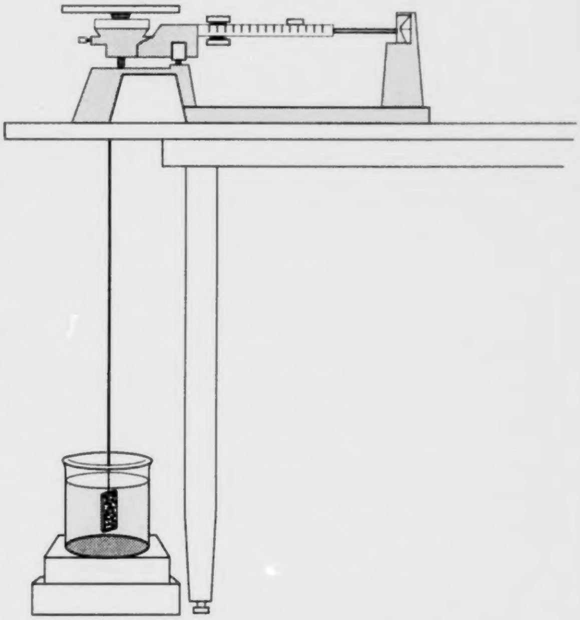


Figure 2. Archimedes' weight displacement apparatus.

Sample No.	1	2	3	4	5	6	7	8
Mass of object (in air)								
Mass of object (in water)								
Mass in air-Mass in water								
Mass of water displaced								
Volume of water displaced								
Volume of object displacing water								

Figure 3. Data table for weight displacement method.

7. Make a graph of the mass of the object in air *vs.* the volume of the object. Record mass along the labelled vertical axis and volume along the labelled horizontal axis. Plot all 6-8 data points on one graph. Give the graph a title; follow accepted graphing procedures.
8. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Is the origin (0,0) a valid point? Explain. Draw one straight line or several straight lines that best "fit" the data.
2. How many lines were plotted on your graph? Predict the number of metals that were used in this experiment, and justify your prediction.

3. The formula for a straight line is $y = mx + b$. What is the value (number reading) of b on your graph? If there is more than one line, do the values of b differ? Since $b = \_$, then $y = \underline{?}$.
4. Using a point on the line, not an experimental point, substitute into the amended equation, mass for y and volume for x, and then solve for m, the slope of the line.
5. Identify a physical property equal to the solution for m in Question 4.
6. Draw a perpendicular line on your graph from the 1-mL mark on the x-axis until it intersects the plotted lines. From the point of intersection, draw a line that is perpendicular to the y-axis. What is the density represented by each slope?
7. Use a reference table to determine the densities of the metal samples. Compare these densities to those determined from reading your graph. Decide which density determination is more accurate, and justify your choice.
8. Explain any discrepancies between density as determined from the table and the graph.
9. A liquid sample that occupies 12.6 mL has a mass of 11.2 g. Find the density of the sample.
10. An Aztec amulet claimed to be pure gold weighs 29.03 g in air and 27.43 g in water. Is the amulet pure gold? The density of gold is 19.3 g/cm^3.
11. In problem 10, if the masses were measured to the nearest tenth of one gram and were 29.0 and 27.4, respectively, does this change the answer?

Implications and Applications

Density, the ratio of mass to volume, is an intensive property. It is often used as a means of checking the identity of a liquid or solid, or checking the concentration of a solution. Jewelry may be checked for its gold content by measuring the density of the specimen. The concentration of antifreeze coolant in an automobile is checked by measuring its density. In many quality control situations, properties are measured and checked to see if the result is within a standard. In some quality control situations, measuring the mass and volume of a sample of product and checking to see if that data pair falls on or near a quality control line is a quality assurance technique.

The amount of a person's body fat may be measured by normal weighing and then weighing while submerged under water. Charts are available that correlate body density with percent body fat.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: Density, A Straight Line Function

Major Chemical Concept

Density is an intensive property; that is, it does not depend upon the amount of substance present. Unlike mass which is dependent upon the quantity of material in an object, the density of a cubic centimeter of aluminum is the same as the density of a cubic meter of aluminum. The density of any object can easily be calculated if only its mass and volume are known. Density is defined as the ratio of the mass of a body to its volume. It is also the slope of a mass versus volume plot. There are several ways to determine these quantities of mass and volume. Students will become familiar with methods for finding mass and volume, thus providing them with the necessary data to calculate density.

Level

Introductory, first-year high school chemistry.

Expected Student Background

See *Related Concepts* and *Related Skills*. Students should be proficient at reading volume and mass measuring devices to the proper precision. Their mathematics backgrounds must be sufficiently strong so that they understand the concept of slope of a line and how slope is calculated. Skills related to preparing graphs are essential in this activity (see *A General Resource* module).

Time

One class period (50 min)

Safety

Read the *Safety Considerations* in the *Student Version*. There are no special hazards associated with this activity.

Materials (For 24 students working in pairs)

12 Graduated cylinders, 100-mL or 50-mL
12 Beakers, 250-mL
72-96 Metal cylinders (or appropriately sized bolts and screws)
12 Centigram balances with platforms, if available
12 Meters thread or nylon monofilament (fishing) line
Scissors

Advance Preparation

There are several ways to attach the thread to the metallic object in order to hang it from the balance. A particularly useful method is to use monofilament fishing line instead of thread. Cut approximately 15-cm fishing line and tie the ends together with an overhand knot. It works best if you hold both ends together to make a "double" overhand knot.

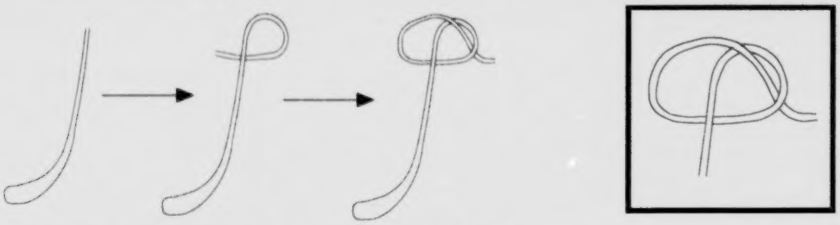


Figure 4. Making a double overhand knot.

Using the loop that you have made, fasten the metallic object by laying the object on the loop and draw one end of the loop through the other end. Draw the loop tightly to the object.

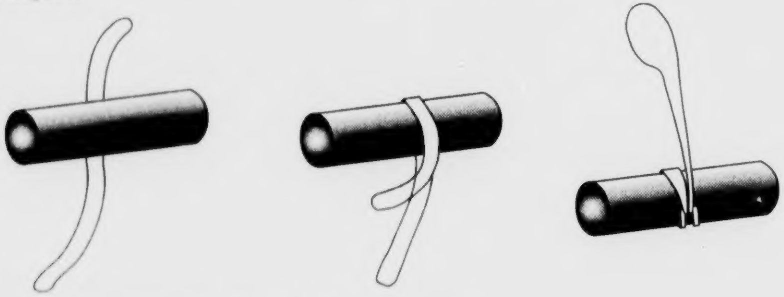


Figure 5. Attaching the loop to the object.

An advantage to using monofilament fishing line is that it doesn't absorb water when submerged under the water's surface. You can use the same loop to measure all the objects whether in the air or in the water. Since the same line can be used for all weighings, the monofilament line can be tared; this increases the accuracy of measurement.

Some student groups could be assigned to do the activity by volume displacement and others by the weight displacement method to compare results. The Archimedes' weight displacement method should provide better accuracy in volume measurement. Using both methods can lead to a discussion on experimental design and precision of measurement. Using the weight displacement method also can lead to a discussion of specific gravity. To illustrate specific gravity, construct a graph labelling the y-axis as volume and labelling the x-axis as (mass in air-mass in water).

If Al and Fe metal cylinders are not available, then Al and Fe bolts and screws of various sizes can be purchased in a hardware store. The volumes should range from 2-8 cm^3. The metallic objects should be divided into two groups of 3-4 of each type of metal. Each metal type should have a full range in the volumes of the metallic objects.

Pre-Laboratory Discussion

This laboratory activity provides an opportunity for students to develop a formula from data collected in an experiment. High school mathematics classes often follow a protocol. First a formula is given. Then a table is constructed using the results of calculations from the formula. Finally, a graph is prepared from the table.

In science, the process is different. First, data are collected (often measurements) and organized into a table. Next, the data are plotted and frequently a line is drawn that "best fits" all the plotted points. Finally, the scientist tries to make sense of the data or interpret the graph. In some instances, the scientist may derive an equation from the graph. In this laboratory activity you will try to develop a mathematical formula from collected data.

The property that will be studied has a constant value that corresponds to the slope of the line on the graph. Since the slope does not change, this property is an intensive property, one whose value does not depend upon the size of the sample.

Teacher-Student Interaction

A goal of this laboratory activity is to provide an opportunity for students to develop a formula from data. The students collect and graph the data. The expected result is a straight line with a zero intercept. Because the line should intersect the origin, the general mathematical function ($y = mx + b$) is reduced to $y = mx$. By substituting the axis labels, it becomes apparent that $y = mx$ can be rewritten as:

$$\text{Mass} = \text{Slope}(m) \times \text{Volume}$$

$$\text{Mass/ volume} = m = \text{Slope}$$

The amount of mass packed into the volume is shown mathematically by the slope and corresponds to the physical property density. The formula is $D = \frac{m}{V}$. Students may encounter difficulty with the mathematical relationship and it might be necessary for the teacher to provide a classroom discussion.

This laboratory activity can be readily transformed into a specific gravity exercise using a weight displacement technique. Plot the mass in air versus the mass difference for specific gravity.

The activity can make use of the volume displacement technique. Students can weigh the dried metallic objects on the balance pan. (Caution students not to drop the metallic object into a glass graduated cylinder since the graduated cylinder might break.) If time is short, have students measure six objects instead of eight. If you choose the weight displacement method, demonstrate how to tare the thread.

Many hardware items such as bolts and screws are not made of pure iron but are alloys of iron. In order to achieve a straight line, select objects known to have the same density—items manufactured from the same material. The physics teacher may have iron cylinders of various sizes that you may use. In some schools the metals (industrial arts) teacher may be able to assist you in making variously sized iron objects.

If possible, have a computer set up in the room so students can record their data and perform the calculations. The data table provided in this exercise can be modified to include density. In addition the computer can graph the data.

This activity also provides an opportunity to discuss extensive and intensive properties, closest packing of ions or atoms in solids, and density and specific gravity problems.

Anticipated Student Results

Sample No.	1	2	3	4	5	6	7	8
Mass of object (in air) (g)	17.78	4.08	27.07	5.40	73.08	37.52	48.08	30.90
Init. volume (of water) (mL)	35.0	35.0	35.0	35.0	50.0	50.0	50.0	50.0
Final volume (of water) (mL)	41.8	36.6	45.0	37.9	59.5	55.0	56.0	54.0
Initial-final volume (mL)	6.8	1.6	10.0	1.9	9.5	5.0	6.0	4.0
Volume of metal object (mL)	6.8	1.6	10.0	1.9	9.5	5.0	6.0	4.0

Figure 6. Representative data table for volume displacement method.

Sample No.	1	2	3	4	5	6	7	8
Mass of object (in air) (g)	17.78	4.08	27.07	5.40	73.08	37.52	48.08	30.90
Mass of object (in water) (g)	11.31	2.56	17.04	3.38	63.69	32.72	41.87	26.91
Change in mass (g)	6.47	1.52	10.03	2.02	9.39	4.80	6.21	3.99
Volume of water displaced (mL)	6.47	1.52	10.03	2.02	9.39	4.80	6.21	3.99
Volume of metal object (mL)	6.47	1.52	10.03	2.02	9.39	4.80	6.21	3.99

Figure 7. Representative data for Archimedes' weight displacement method.

Answers to Data Analysis and Concept Development

1. Yes, zero volume would also have zero mass. (You can also argue for the answer no, since the graph should include only measured data. Using the intercept as a data point biases the measured data.)
2. There are two straight lines indicating two kinds of metals were used. Each line has a slope that is dependent upon the mass/volume ratio. This same ratio is a physical constant for a particular metal.
3. $b = 0$, therefore $y = mx$
4. $m = \frac{y}{x} = \frac{\text{mass}}{\text{volume}}$
5. Density
6. Answers will vary. The density of aluminum is 2.70 g/cm^3 and the density of iron is 7.86 g/cm^3. Student answers from reading the graph will vary within a range near the accepted values.
7. Density values determined from the data table should vary in a range more closely to the accepted values than from using the slope to determine the density values.
8. The density value read from the graph is probably less accurate inasmuch as the line is a "best fit" line. Using all of the data points and a statistical procedure, like the method of least squares, to calculate the slope would lead to a better result from the slope as compared with the tabular data. Measuring volumes from calibrated cylinders presents the largest errors. Steel is an alloy, and measuring the density of steel gives a result different from that expected for iron.
9. $\frac{11.2\text{ g}}{12.6\text{ mL}} = 0.889\text{ g/mL}$
10. Mass in air – mass in water = mass of water displaced; 29.03 g – 27.43 g = 1.60 g = mass of water displaced.

    Since the density of water is 1.00 g/mL, the *mass* of water displaced is equal to the *volume* of water displaced. This in turn is equivalent to the volume of the object displacing water. Density $= \frac{29.03\text{ g}}{1.60\text{ cm}^3} = 18.1\frac{\text{g}}{\text{cm}^3}$. The amulet is not pure gold. See the October 1989 issue of *ChemMatters* (page 16) "Indiana's Error" for an interesting story dealing with the density of gold.

11. When the masses are measured to this level of accuracy, the minimum mass is 28.9 g in air and the maximum mass is 27.5 g in water. This density is 21 g/cm^3. So, unless the masses are measured with sufficient accuracy that the difference will have more than two significant figures, answering the question is impossible (see *General Resource* module for discussion of uncertainty in measurement and significant figures).

Post-Laboratory Activities

1. Discuss the student results. Compare densities obtained by the volume displacement method to densities obtained by the Archimedes' weight displacement method. Have students locate the accepted values for the densities of these metals in *The Handbook of Chemistry and Physics* or in a chemistry or physics text.
2. Reemphasize the significance of the slope of a mass versus volume curve in determining the density of a substance. Ask the class to compare the degree of slope of the aluminum line and the iron line to the calculated densities of these metals.
3. For an extension activity, ask students to determine a way to find the density of a material less dense than water (will float on water). Cork stoppers found in your chemistry laboratory would be good candidates for this exercise.

Assessing Laboratory Learning

1. The figure below shows a graph of mass *vs.* volume for two immiscible liquids. Predict from the graph what will happen if Liquid A and Liquid B are poured together. *[The liquids are immiscible and thus will not mix but will separate with Liquid A settling below Liquid B in the container. The reason for this occurrence is that Liquid A has a greater density than Liquid B. The density is indicated by the slope of the line plotted on a mass* vs. *volume graph.]*

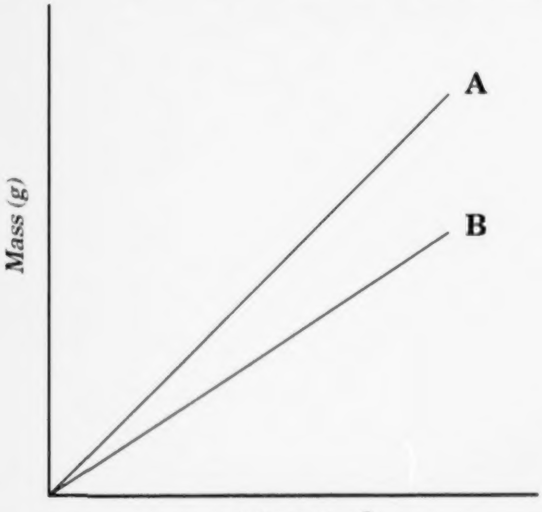


Figure 8. Plot of mass vs. volume for two immiscible liquids.

2. Ask students to explain the principles of the conventional battery tester or antifreeze tester. *[Both battery testers and antifreeze checkers use large bulb pipets (similar to kitchen meat basters) to withdraw a sample of fluid from the battery or radiator. The stem of the pipet contains color-coded beads of known density. By observing which beads float and which sink in the liquid withdrawn, the gas station attendant can estimate the concentration of the acid or antifreeze in the liquid. This rapid, semiquantitative procedure is more than adequate for the purposes of deciding what course of action to take in automobile maintenance.]*

LABORATORY ACTIVITY: STUDENT VERSION

Activity 2: Energy Needed to Melt Ice

Introduction

Ice at 0 °C is added to a measured volume of warm water. Some of the ice melts. Once the mixture reaches 0 °C, the volume of liquid water is measured. The additional volume comes from melted ice. Using the amount of ice melted together with the temperature and volume of warm water, the heat of fusion of the ice is determined.

Purpose

Students will determine the energy needed to melt one gram of ice by measuring the heat lost by a known quantity of warm water when ice is placed in the water and melted.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. A burner will be used to warm the water to about 50 °C. Burns are possible. Avoid contact with heated objects. Use hot pads to move warm objects. Always place warm objects where someone else will not inadvertently touch them.

Procedure

Work in pairs to accomplish the following steps.

1. Warm water to about 50 °C. Add about 100 mL to a graduated cylinder. Record the volume.
2. Pour the water into a Styrofoam® cup. Measure and record the temperature of the warm water to the nearest 0.2 °C.
3. Obtain several ice cubes; pat dry with a towel and immerse in the cup.
4. Cover with a lid that has a small hole near the edge. Allow to equilibrate at 0 °C by shaking and adding ice until the system is at 0 °C and excess ice remains.
5. Pour liquid water into a beaker, holding back ice that reamins.
6. Measure the volume of the liquid by pouring water from the beaker into a large graduated cylinder. Record the volume.
7. Repeat Steps 1 to 6 as instructed by your teacher.
8. The sample data table should be transferred into your laboratory report.
9. Clean up your laboratory station and place the equipment in its proper location.
10. Thoroughly wash your hands before leaving the laboratory.

	Trial I	Trial II
Volume of warm water used		
Initial temperature of water		
Equilibrium temperature of ice/water system		
Volume of water, final		
Mass of ice melted		
Change in temperature of water		
Energy released by water as it cools		
Calories required to melt 1.0 g of ice		
Kilojoules required to melt 1.0 mol of ice		
Percentage error (experimental *vs.* accepted)		

Figure 9. Sample data table.

Data Analysis and Concept Development

Given the sample data table:

1. Determine the mass of ice melted (the density of water is one gram per milliliter).
2. Determine the change in temperature of water.
3. Calculate the energy released by the liquid water as it cooled.
4. Calculate the calories of energy required to melt one gram of ice? The specific heat of water is 1.0 cal/g °C.
5. Using the results of this experiment, determine the number of kilojoules required to melt one mole of ice.
6. Use chemical symbols in writing an equation for the fusion (melting) of ice, and include the energy term (in kJ /mole of water) as calculated in Question 5.
7. Calculate the experimental (percentage) error by comparing the experimental value which you have determined for the heat of fusion of ice to the accepted value for the heat of fusion. *[6.010 kJ/mol.]*

Implications and Applications

There are innumerable situations in which we try to control the temperature of foods and other materials using ice. Ice is very good for keeping materials cool in warm surroundings because is requires so much heat to melt.

When traveling in frigid locales, automobile drivers are encouraged to include kits with blankets, high energy foods that don't perish (such as peanut butter), and also to carry a metal cup and a candle. The cup and candle are intended to prepare emergency water by melting snow. It is dangerous and perhaps impossible to use body heat for this purpose when stranded in subfreezing weather.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 2: Energy Needed to Melt Ice

Major Chemical Concept

Heat flows from warmer to cooler bodies. In a closed system, the heat energy lost by the warmer body must equal the heat energy gained by the cooler body. The Styrofoam® cup reduces the heat flow between the water and the outside room to a minimum. The volume inside the cup can be considered to be a closed system where there is no inflow or outflow of heat or matter. Hence, the heat energy exchange can be considered to be only between the warm water in the cup and the ice which the students are adding to the cup.

By knowing the mass of warm water and the temperature change which it undergoes, one can determine the heat lost by the warm water when the ice is added. This is the same quantity of heat which is overcoming the attractions in the ice crystal and causing the ice to melt. The mass of ice melted can be determined. Since the mass of ice melted and the heat energy used to melt it are known, the molar heat of fusion is just a quick calculation away.

Level

Introductory, first-year high school chemistry.

Expected Student Background

See *Related Concepts* and *Related Skills*. Students must understand the concept of heat flow from warmer to cooler bodies. They must also be somewhat familiar with the concept of a "closed system." It would be very helpful for the mathematical work in this laboratory if the students are proficient in direct proportion equations and the analysis of units (labels) on measurements.

Time

Teacher preparation: 15 min

Student laboratory time: Approximately 30 min per trial.

Safety

Read the *Safety Considerations* in the *Student Version*. Caution students to be careful with the heated objects. The burner, ring, and wire gauze are all possible sources of burns. Caution students not to allow chunks of ice to drop on the floor; slipping on ice chunks might lead to serious injury.

Materials (For 24 students working in pairs)

36 Beakers, 250-mL
12 Styrofoam® cups, 12-oz, with lid
12 Burners, support stands, rings, and wire gauze
12 Graduated cylinders, 100-mL
12 Graduated cylinders, 250-mL
12 Thermometers, –10 °C to 110 °C
Ice cubes
Paper towels
Hot pad

Advance Preparation

1. The ice cubes should be wet, that is, at melting temperature. They should not be used immediately after being removed from the refrigerator. Allow them to stand at room temperature a few minutes. It is instructive for many students to use a thermometer to measure the temperature of ice that has been freshly removed from a deep freeze. Many students do not believe that the temperature of ice can go below 0 °C.
2. Styrofoam® ice chests are very efficient in keeping ice throughout the day.

Pre-Laboratory Discussion

Even though the melting of ice and freezing of water are common occurrences, the concept of heat released during freezing seems counterintuitive. Citrus fruit growers spray their orchards during freeze warnings to prevent damage to the fruit trees. The environment is affected by the phase change involving water and ice. A melting snow cover will prevent the temperature from rising as high as it would without snow and the freezing of snow melt after sunset prevents the temperature from dropping as rapidly as it would otherwise.

In this particular activity we will determine the heat needed to melt 1.0 g of ice. This is the same amount of heat released to the surroundings during freezing.

Teacher-Student Interaction

During the session it is important to circulate and ask questions about freezing and thawing phenomena.

Anticipated Student Results

	Trial I
Volume of warm water used	98.7 mL
Initial temperature of water	46.6 °C
Equilibrium temperature of ice/water system	0.0 °C
Volume of water, final	158.9 mL
Mass of ice melted	60.2 g
Change in temperature of water	–46.6 °C
Energy released by water as it cools	4.60×10^{3}cal
Calories required to melt 1.0 g of ice	76.4 cal/g
Kilojoules required to melt 1.0 mol of ice	5.75 kJ/mol
Percentage error (experimental *vs.* accepted)	4.4%

Figure 10. Sample data table.

Answers to Data Analysis and Concept Development

1. $158.9\text{ mL} - 98.7\text{ mL} = 60.2\text{ mL}$

   $60.0\text{ mL} \times 1.00\frac{\text{g}}{\text{mL}} = 60.2\text{ g}$

2. ΔT = final temperature - initial temperature $= T_f - T_i$

   $= 0.0\ ^\circ\text{C} - 46.60\ ^\circ\text{C}$

   $= -46.6\ ^\circ\text{C}$

3. $\text{Energy lost} = \Delta T \times 98.7\text{ g} \times 1.00\frac{\text{cal}}{\text{g}\ ^\circ\text{C}}$

   $= -46.6\ ^\circ\text{C} \times 98.7\text{ g} \times 1.00\frac{\text{cal}}{\text{g}\ ^\circ\text{C}}$ (specific heat of water)

   $= -4.60 \times 10^3\text{ cal}$

4. $\text{Energy} = \frac{4.60 \times 10^3\text{ cal}}{60.2\text{ g of }H_2O} = 76.4\text{ cal/g}$

5. $76.4\frac{\text{cal}}{\text{g}} \times 18.0\frac{\text{g}}{\text{mol}} = 1.37 \times 10^3\frac{\text{cal}}{\text{mol}}$

   $1.37 \times 10^3\frac{\text{cal}}{\text{mol}} \times 4.184\frac{\text{J}}{\text{cal}} \times \frac{1\text{ kJ}}{10^3\text{ J}} = 5.75\frac{\text{kJ}}{\text{mol}}$

6. $H_2O(s) + 5.75\text{ kJ} \rightarrow H_2O(l)$

7. $\%\text{ error} = \frac{|\text{Experimental - Accepted value}|}{\text{Accepted value}} \times 100\%$

   $= \frac{|5.75\text{ kJ} - 6.010\text{ kJ/mol}|}{6.010\text{ kJ/mol}} \times 100\%$

   $= 4.3\%$

Post-Laboratory Discussion

1. Compare student answers. Have students write their answers and percentage error on the chalk board for later discussion or have the results placed into a spreadsheet on the class computer.

2. Have a student look up the heat of vaporization of water. Lead a discussion of what this heat of vaporization indicates about the attractive forces in liquid water compared to the attractive forces in solid water. *[Students will find the heat of vaporization to be 40.68 kJ / mole for water. The ratio of the heat of vaporization to the heat of fusion is about 6.75:1. The attractive forces in the liquid state must be stronger than the attractive forces in the solid state as much more energy is required to separate the liquid molecules into gas molecules.]*

Assessing Laboratory Learning

1. What would be the effect of replacing the Styrofoam® cup with a cup made of thin copper metal? *[The system would no longer be a closed system of only the warm water and the ice. The copper cup is an excellent conductor of heat (unlike Styrofoam®) and heat flow will continue into or out of the copper cup into the warm water and ice, according to whether the temperature is above or below the air temperature.]*
2. Why is it important that there be ice in the system at the end of the experiment, in other words, why is it important to have the system at 0 °C at the end of the experiment. *[If the temperature moves above 0 °C then the ice has all melted and some of the heat energy from the hot water has been used to actually warm the cold water which came from the ice. The arithmetic problem of solving this system is somewhat more complicated than for the case used.]*

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Which Will Evaporate First?

Purpose

To observe the rate of evaporation of various volatile liquids and to discuss the results in terms of intermolecular forces and vapor pressure.

Materials

Samples of the following volatile liquids:
- Methanol
- Ethanol
- 1-Propanol
- 2-Propanol
- 1-Butanol
- Acetone
- Water
- Watch glass for each liquid
- Cotton ball for each liquid
- Pair of disposable gloves
- *Very clean* chalkboard

Safety and Disposal

All of these liquids except water are flammable and their vapor-air mixtures are explosive. Do not have any flames or sparks of any kind in the room during this demonstration. The vapors of 1-butanol and 2-propanol are irritating to the eyes and respiratory system. Do not over use. Any remaining liquid in the watch glasses may be washed down the drain with a 20-fold excess volume of water. The moistened cotton balls should be placed into a container of water, squeezed out, and thrown into a waste can. The water in the container can then be flushed down the drain.

Procedure

1. Write the names of the selected volatile liquids on the clean chalkboard. Have the students help you to write under the names the formulas and structures of the liquids.

2. Start with methanol and ethanol. Have the students predict which liquid will evaporate first. Use plastic gloves. Place a cotton ball in the watch glass containing these two liquids, squeeze off excess, and make a swipe down the chalkboard with each liquid under its name. Measure the time necessary for each to evaporate.
3. Check the students' predictions and discuss their reasoning.
4. Repeat with 1-propanol and 2-propanol; with 1-butanol and acetone; with methanol and water. Be prepared to repeat with pairs of liquids from different sets to support or disprove suggested hypotheses.

Remarks

Volatility of liquids is dependent upon the strength of the intermolecular forces. Discussion of London dispersion forces, dipole-dipole interaction, and hydrogen bonding should be utilized. The effect of structural isomerism on the degree of polarity is nicely illustrated.

Liquid	Boiling point (°C) (at 760 mmHg)	Vapor pressure (mmHg) (at 298 K)	Heat of vaporization (kJ/mol)
Acetone	56.5	220	32.0
Methanol	64.7	123	39.2
Ethanol	78.4	56.0	40.5
Water	100	23.8	40.7
2-Propanol	82.5	42.9	42.3
1-Butanol	118	7.27	46.0
1-Propanol	97.8	19.8	47.3

Figure 11. Boiling points, vapor pressures, and heats of vaporization for selected liquids.

Reference

This demonstration is modified from *Experiments and Demonstrations* for the 1986 ICE Workshop in Tucson.

Demonstration 2: Boiling Water the Fast Way

Description

Warm water is drawn into a large syringe, and the end is sealed. When the plunger of the syringe is drawn out and the pressure drops, the warm water quickly boils.

Materials

Plastic syringe, 30-mL or 50-mL
Cap for the syringe
Beaker of warm water

Hazards

If the water is too warm, a skin burn may result.

Procedure

1. Place open end of syringe into the beaker of warm water. The water does not have to be extremely warm; 50°C to 60°C suffices. In fact, the lower the temperature at which the water can be made to boil, the more impressed students will be.
2. Draw approximately 10-mL water into the syringe.
3. Place the plastic cap on the end of the syringe if one is available. If no cap is available to seal the syringe, place the end of your index finger over the end of the syringe or plug the end with modeling clay.
4. Pull on the plunger until it is nearly out of the syringe. Do not pull so hard as to "pop" the plunger completely out of the syringe. The water should begin to boil. If it does not, try warmer water.
5. Let the plunger slide back into the syringe. The boiling will stop. Repeat this process as often as you like. You will need to use a fresh sample of warm water once boiling cannot be sustained. It is possible to get the plunger to stay out by drilling a hole in the stem of the plunger and inserting a nail in the hole as the plunger is extended. This will help to keep your fingers from getting tired.

Demonstration 3: Relieving the Tension

Purpose

To demonstrate surface effects.

Materials

1-L Beaker or shallow glass pie pan
1 g Flowers of sulfur (fine sulfur powder)
1 mL Dishwashing detergent in beaker
Tap water
Petri dish
25 mL Homogenized fresh whole milk (4% milk)
Food coloring
Thin-stem transfer pipet
Triple-beam balance
Support stand with ring
Wire circle with wire loop (60 cm heavy gauge aluminum wire; solder; soldering iron; suitable soldering work surface)

Safety and Disposal

There are no unusual hazards in this experiment. No special precautions are required in this experiment; follow routine laboratory precautions. The materials used in this activity may be disposed of safely at the sink.

Procedure

1. Place 500 mL tap water in a 1-L beaker. Sprinkle some powdered sulfur on the surface of the water. Dip your clean dry finger into the water, and note the effect on the sulfur.
2. Dip your finger into some liquid dish detergent, and then dip it into the beaker containing the sulfur. Note the effect on the sulfur.
3. Fill a Petri dish bottom half-full of whole milk. Drop some food coloring onto the surface. Fill a transfer pipet with some liquid dishwashing detergent. Place a few drops of the detergent on the inside wall of the Petri dish above the level of the milk. Note any evidence for reaction. After a long period of time, the apparent changes cease.

4. A flat wire circle with a loop is constructed for this part of the experiment. Hang the loop from the balance hook of a triple beam balance. (If the balance has no hook, construct one using wire.)

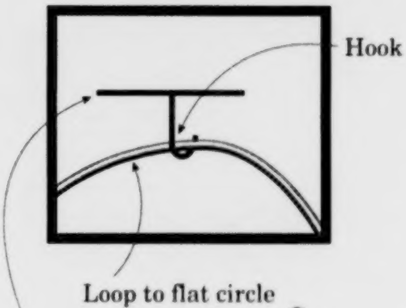


Figure 12. Circle with loop.

5. Fill a Petri dish bottom half-full of distilled water. Add a drop of food coloring to make the water more visible. Support the Petri dish on the ring of a support stand. Bend the wire loop so that the plane of the flat wire circle is precisely parallel to the plane of the surface of the liquid. Adjust the level of the dish so that the ring is just below the surface of the liquid in the dish.
6. Add weight by carefully sliding weights on the balance toward greater values.

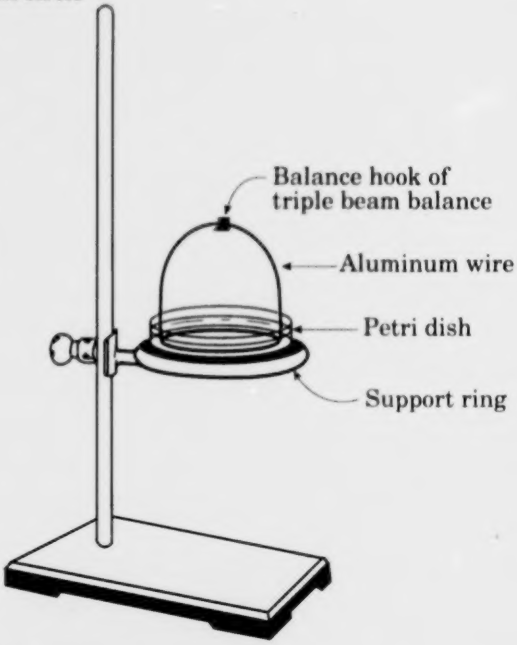


Figure 13. Apparatus set up.

Remarks

The sulfur must be in the form of a very fine powder. Sulfur floats on water. The surface tension of the water is high, and the water is unable to wet the sulfur. The density of the sulfur is greater than that of the water. A small amount of detergent, just enough to go unnoticed on fingertips, will sufficiently lower the surface tension of the water to allow it to wet the sulfur, which immediately sinks.

A detergent (surfactant) changes the surface characteristics of a suspension. While the phenomena involved are very complex, changes in surface tension play a major role in the mixing effect seen on changes in food coloring puddled on the surface of some milk once a detergent is added to the surface.

The surface tension in water is sufficient that nearly 1 cm of film can be drawn from a water surface using a suitable device such as a flat wire ring and a gentle steady source of pull such as that from a triple beam balance.

To manufacture a ring for the students, bend a smooth, flat circle of heavy gauge aluminum wire into a circle of 8-10 cm diameter. Once the wire lies flat and the loop closes with a slight spring-like pressure, solder the circle closed. Fashion a loop of the same wire with a span equal to the diameter of the wire circle. Solder this loop perpendicular to the circle.

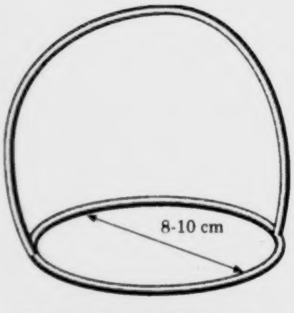


Figure 14. Flat wire ring.

Using the balance to lift the ring from the water's surface is the most difficult part of this experiment.

Reference

(See Experiment C33 of *Doing Chemistry*.)

Demonstration 4: Slow the Flow

Purpose

To illustrate the effect of viscosity upon rate of liquid flow.

Materials

4 Graduated cylinders,100-mL
SAE 10- and 40-weight motor oil (or alternate with low and high SAE numbers)
2 Beakers,1000-mL
Hot plate or burner
Ice

Safety and Disposal

Hot water may cause burns. Retrieve the motor oil to be stored and used again next year. Wash out the cylinders with detergent and dispose of the waste in the sink.

Procedure

1. Add 100 mL SAE 40-weight oil to two of the graduated cylinders.
2. Place one of these cylinders into a beaker containing hot water and the other into a beaker containing cold water. Allow them to stand for about five minutes. Perform Step 3 while waiting.
3. Add 100 mL SAE 10-weight oil to the third graduated cylinder and 100 mL SAE 40-weight oil to the fourth graduated cylinder.
4. Drop a plastic bead or glass bead into each cylinder one at a time. Record the time it takes for the bead to sink to the bottom in each cylinder.
5. Remove the first two graduated cylinders from their respective water baths. Predict the times needed for a bead to fall to the bottom of the cylinder. Drop the beads and measure the duration of the fall to the bottom.

Remarks

Ask students the following:

1. On a cold winter day in Fargo, ND, which motor oil would you want to use to start your car? *[SAE 10 for easy low-temperature starting.]*
2. During hot weather, which motor oil provides the most protection to your car? *[SAE 40 for better load capacity in bearings at normal running temperature.]*
3. Distinguish between density and viscosity? *[Density is a property of matter representing mass per unit volume. Viscosity is the resistance of a liquid to flow.]*

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. Solids and liquids occupy roughly 1/1000th the volume of the corresponding gases. Determine the ratio of the distance between solid or liquid particles to the corresponding distance between gaseous particles? *[The solid and liquids are about 10 times closer together. Volume is length cubed. Particles moving 10 times closer in all of their linear dimensions actually move into 1 / 1000th as much volume.]*

2. State the physical evidence that supports the idea that particles in both solids and liquids are very close together. *[Neither solids nor liquids are very compressible, hence there is little space between their particles. Gases, on the other hand, are relatively easily compressed.]*

3. Criticize the statement "Water boils at 100 °C." *[Water only boils at its "normal" boiling point if the sample is pure water under one atmosphere of pressure. Atmospheric pressure variation and/or impurities in the water cause the boiling point to differ from the norm.]*

4. Describe similarities and differences among the warming curves of different crystalline substances. *[All warming curves of crystalline materials have portions sloped upward to the right, indicating the speeding up of solid, liquid, or gaseous particles (increasing kinetic energy). All have flat portions, indicating change of phase of the particular substances (increasing potential energy). The slopes for individual materials will vary, indicating differences in energy necessary to speed up the particles (specific heat). The lengths of the flat portions will vary, indicating differences in the amount of energy necessary to overcome attractive forces within the substances. The flat portions will occur at different temperatures, indicating that substances have different melting and boiling points.]*

5. State the temperature effect observed during the addition of heat to a substance at its melting point. *[Until all solid melts, the temperature remains the same. At the melting point, added heat is utilized to separate particles and increase potential energy rather than to speed up the particles of the substance.]*

6. Many scientists believe that earth's inner core is solid while its outer core is liquid. State conditions that might make this possible even though the inner core is thought to be at a higher temperature than the outer core? *[The inner core is under tremendous pressure. The particles in the inner core are pressed so tightly together that they are bonded into a solid crystal lattice even though the kinetic energy is such that the material would be liquid at lower pressure.]*

7. Suppose an insulated container held a solid/liquid mixture of a certain substance at its melting point so that heat neither entered nor left the container. Predict what would be observed. *[Melting and freezing would continue at equal rates. A dynamic equilibrium would be established. The heat needed to melt the substance would be released as an equal amount of the substance would freeze. The amount of solid and liquid would remain constant even though melting and freezing would continue unabated. (Although the size of the solid "chunk" would not change, its shape would change over time.)]*

8. Explain the loss of solid and liquid lubricants in devices sent into outer space. *[Even though liquid and solid greases have extremely low vapor pressures, the prevailing pressure in outer space is below this pressure, and thus the lubricant will boil away violently.]*

9. Compare the distance between solid and liquid particles. *[The distances between liquid particles and solid particles are nearly the same: both liquids and solids are essentially incompressible, and there is little expansion as substances go from the solid to liquid phase.]*

10. Why are liquids, but not solids or gases, utilized in hydraulic systems such as automobile brakes? *[Solids aren't fluid and cannot flow throughout the hydraulic system. Gases are fluid and can flow throughout the hydraulic system, but since gases are extremely compressible, they cannot transfer the pressure.]*

11. At the melting point of a substance, it consists of a mixture of solid and liquid portions. State the molecular differences expected between the solid and liquid portions. *[The solid particles must be distinguished from liquid particles by observing the differences in their motions. Since both the solid and liquid particles are at the same temperature, they have the same average kinetic energy but may be distinguished by differences in the amount of translational, rotational, and lattice vibrational kinetic energy. The liquid particles possess potential energy because they are farther apart when compared to particles in the solid phase.]*
12. Motor oil has a low vapor pressure compared to water. Compare the intermolecular forces in these materials. Predict which liquid has the higher boiling point. *[The fact that motor oil has a low vapor pressure indicates that the attractive forces between the relatively large oil molecules are quite great. Few of these molecules escape easily into the gaseous state, thus the low vapor pressure. Since the vapor pressure is low for motor oil, one would have to heat the oil to a high temperature to increase its vapor pressure to atmospheric pressure. Oil boils at a higher temperature than water.]*
13. It is possible to place warm water in a syringe and to cause the water to boil by pulling out on the plunger of the syringe. Explain the physical changes that make this possible. *[As the plunger is drawn out, pressure in the syringe drops low enough that it is below the vapor pressure of the warm water. The water begins to boil.]*
14. Which is more dense, milk or cream? Support your answer. *[Milk is more dense than cream; cream floats on milk. (Students frequently confuse cream's higher viscosity with density.)]*
15. Nurseryworkers "mist" their bedding plants with water on cold spring nights when frost is predicted. Describe the process that protects the plants. *[As the misting water cools and freezes, a great deal of heat is released into the surrounding atmosphere. This heat keeps the ambient temperature higher than it would be without the cooling/freezing water, thus protecting the tender plants.]*
16. Describe the molecular changes that cause the potential energy of a substance to increase when that substance is melted or boiled? *[The increase in potential energy at the melting or boiling point is caused by overcoming attractive forces and moving particles farther apart. Gases have the potential for attracting one another and moving closer together thus decreasing the amount of potential energy between the particles as the substance liquifies. This decrease in potential energy manifests itself as an increase in heat flow out of the system.]*
17. Describe the molecular cause of the expansion characteristics of water near its freezing point? *[Hydrogen bonding in the polar water molecules produce an open hexagonal crystal lattice causing water to expand during the phase change.*

Counterintuitive Examples and Discrepant Events

Boiling Water with Ice

Materials

Flask, round bottom, 500-mL
Rubber stopper for flask, solid
Ring stand
Ring of appropriate size to hold inverted flask
2 or 3 Ice cubes
1 Oven mitt to fit hand
Safety goggles
Hot plate or burner

Safety and disposal

A round bottom flask is less likely to break under the reduced pressure produced in this demonstration. The boiling water and steam will cause burns. Wear the oven mitt for protection. At the end of the demonstration, gently pry the stopper off and pour the water down the sink. The stopper will be somewhat difficult to remove as the pressure on the outside of the flask is substantially greater than the pressure within the flask.

Procedure

Place 200 mL of water in the 500-mL round-bottom flask. Heat the water to a vigorous boil in full view of the class. Using the oven mitt, remove the flask and stopper it immediately. Let the water cool for (at most) a minute. Hold the flask with the mitt. Invert the flask into the ring on the ring stand. Have a student come up in front of the class. Have the student put on safety goggles and then take an ice cube in each hand and rub them on the rounded bottom of the flask. The water will soon (15 to 30 seconds) begin to boil vigorously. The demonstration can be repeated for up to 10 minutes.

Background

Placing a stopper on the flask shortly after the boiling water is removed from the heat source, most of the "air space" above the water is filled with water vapor, not air. Ice placed on the top of this space causes water vapor at (relatively) high temperature to condense. When removed from the vapor state, the contribution of water vapor to the total pressure inside the flask is lowered. When the total pressure inside the flask goes below the vapor pressure of water (the warm water), that water begins to boil. So, even though the water is considerably below the normal boiling temperature, it does boil at the ambient pressure inside the flask.

Reference

Taken from *Doing Chemistry*, Experiment D24.

Floating and Sinking Ice

Students expect an ice cube to float on water. They know that ice is less dense than water. In this demonstration, an ice cube will float in one container of water and sink in another container of "water."

Materials

2 Beakers, 250-mL
150 mL Ethanol or isopropyl alcohol
150 mL Water
2 Ice cubes

Safety and Disposal

Have no open flames; the alcohols are flammable. Flush ethanol or isopropyl alcohol down the sink with ten volumes of water per volume of alcohol.

Procedure

1. Pour 150 mL water into one beaker and 150 mL ethanol into the other beaker prior to class. Position the two beakers where all can see and place an ice cube into each beaker. Ask the students to write an explanation as to why the cube sinks in one beaker and floats in the other beaker.
2. Allow them to ask two or three questions about this system which can be answered "Yes" or "No" and then give them time to write a new explanation if they wish. Choose students to read their explanations.

3. When a group consensus is reached that the reason for the sink/float discrepancy is because of the differences in the densities of the two liquids (water and alcohol), ask them what should happen when a piece of steel is placed on the water. Since steel is more dense than water, the steel should sink. Carefully place a straight pin on the surface of the water in the beaker. It will float. Have them explain why this more dense substance floats on water. This will provide a good lead-in to a discussion of surface tension.

Density: "Not on the Level"

We all expect that fluids will "seek their own level." This counterintuitive event will leave students searching for a reason for why the two levels of "water" in a U-shaped tube are not the same. The explanation is that two liquids with different densities have been added to the two sides of the U-shaped tube.

Materials

2 U-shaped glass tubes (about 6-8 mm in diameter and 20 cm or more in height)
10 mL Saturated salt solution (add sodium chloride to 15 mL of water until no more salt dissolves after stirring) or ethanol or isopropyl alcohol
10 mL Water
Food coloring

Safety and Disposal

No special precautions required for fresh and salt water. Eliminate spark sources and open flames if ethanol is utilized. Liquids may be disposed of in the sink. (Flush drain with plenty of water if alcohol is used.)

Procedure

1. Pour fresh water into one U-tube until it is about 2/3 full and note that the water level is the same on both sides. Placing food coloring in the water prior to the demonstration will help students view the liquid levels.
2. Pour salt water into a second U-tube until it is about 1/3 full. Pour fresh water into one side of the second U-tube until it is about 2/3 full. Note that the fresh water side is at a higher level than the salt water side.
3. Ethanol or isopropyl alcohol may be substituted for the salt water with opposite results. Running the trial again and coloring the fresh water and salt water or ethanol differently may help the students to understand what is happening in this demonstration.

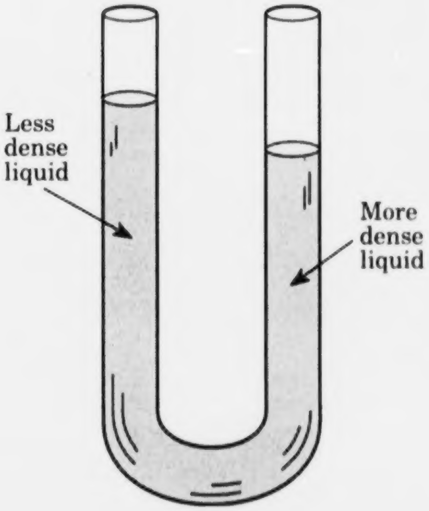


Figure 15. "Not on the level" demonstration.

Remarks

Have the students write a short explanation of what they believe is causing the discrepancy in liquid levels in the second U-tube.

Reference

Taken from *Doing Chemistry* A02.

A Non-Newtonian Fluid: Cornstarch Putty

Cornstarch putty belongs to a class of materials which does not obey the usual laws of viscosity. Also included in this class of materials (known as non-Newtonian fluids) are the commercial products, Slime™ and Silly Putty™. Under low stress, such as slow pulling or pouring, these materials flow and stretch much as normal fluids do. But when hit with a hammer, Silly Putty™, for example, loses any resemblance to a fluid and actually shatters as if solid.

Cornstarch putty exhibits similar behavior. If poured from container to container, cornstarch putty looks like pancake batter. If one quickly "slaps" the surface or "grabs" a handful of this material, it becomes a rigid, solid-like material.

Materials

Plastic mixing bowl
500 mL Cornstarch
250 mL Water
Mixing spoon
Pie pan
Safety goggles
Laboratory apron
Paper towels

Safety and Disposal

No special precautions are required. Upon standing the putty will separate but can be reconstituted by kneading with your hands. The cornstarch putty may be disposed of in the laboratory sink with an adequate supply of water.

Procedure

1. Place the 500 mL cornstarch in the plastic mixing bowl.
2. Slowly add the 250 mL water while mixing the cornstarch with the mixing spoon or while kneading by hand. (Check the consistency prior to the demonstration of slapping the mixture in 4 below.)
3. In the presence of students, pour the cornstarch putty from the mixing bowl into the pie pan. It will pour like a normal fluid.
4. Place a laboratory apron and safety goggles on a willing class volunteer. Ask the volunteer to *quickly* slap the surface of the putty in the pie pan. *(The volunteer and the class will expect to have a cornstarch shower. Instead the surface won't even be penetrated by the slap.)*
5. Instruct the student to reach into the pie pan and to "grab" a quick handful of the putty. *(It will come out as a "solid" mass which will almost immediately liquefy and drip off of the student's fingers like pancake batter.)*
6. Supply the student volunteer paper towels to wipe off the cornstarch putty so that it doesn't drip all over the laboratory table.

Reference

Walker, J. (1978). The amateur scientist; serious fun with polyox, silly putty, slime, and other non-Newtonian fluids. *Scientific American, 239*(5), 142-149.

"Melting" Dry Ice

Dry Ice, solid carbon dioxide, is useful because it is very cold and because, under room pressure, it evaporates directly from the solid to the gaseous state leaving no messy liquid residue. (Evaporation from the solid to the gaseous state is called sublimation. Dry Ice sublimes.)

Purpose

To observe the melting of dry ice from solid to liquid at high pressure.

Materials

Safety goggles
Gloves
Beral™ pipet, 4 mL with wide stem (3-5 mm bore) *(use only a polyethylene pipet)*
0.5 lb Dry ice
Scoopula with handle
Towel, hammer, Styrofoam® box
Scissors
Pliers (with flat surface)

Safety and Disposal

Handle the dry ice with care to avoid contact. Wear gloves. The teacher should powder the dry ice. Wrap a 0.5-1.0 lb piece of dry ice in a towel or canvas cloth. Smash the wrapped ice against a hard surface such as a concrete floor. Place the powdered ice in a Styrofoam® or other insulated container. Dispose of excess dry ice by allowing the solid to sublime in a well-ventilated area.

Procedure

1. Use scissors to cut the stem of a wide-bore tip 4-mL transfer pipet (Beral™ pipet) to a length of 3 cm.
2. Wear eye protection. Protect hands with gloves or towel, or use scoop with handle. Repeatedly squeeze the pipet bulb, place it into powdered dry ice, release quickly to draw in powder, and tilt to trap powder until about 1-2 cm^3 of dry ice power are in bulb.
3. Wear eye protection. Use pliers to squeeze the pipet tip shut. Observe.
4. Once the solid melts, release the seal of the pliers on the bulb. Observe.
5. Wear eye protection. Use pliers to squeeze the pipet tip shut. Observe until the pipet ruptures.

Remarks

Use only a polyethylene pipet. As the dry ice sublimes, it builds up pressure. When the triple point pressure is reached inside the bulb, the solid melts to form liquid carbon dioxide. Releasing the grip of the pliers drops the pressure and cooling due to expansion of the gas is sufficient to refreeze much of the remaining carbon dioxide. Holding the bulb sealed usually leads to an increase of gas pressure that exceeds the strength of the pipet and causes it to rupture. This is usually a gentle process as the polymer in the bulb stretches and distorts rather than shatters. Note: If you find that the pipet ruptures before the triple point is reached, try Tygon™ tubing, clamping a 6-in. piece at both ends.

The observed changes can be interpreted in terms of a phase diagram for carbon dioxide. When solid dry ice is exposed to the atmosphere, the solid sublimes and the pressure never exceeds one atmosphere. When the bulb tip is clamped, the pressure can rise. Adding heat to the clamped system causes the temperature of the solid dry ice to rise. When the temperature reaches about −56 °C, the pressure is just over 5 atmospheres and the triple point is reached. Under these conditions, the liquid phase of carbon dioxide is stable. Releasing the pressure permits cooling by expansion. The solid reforms, the pressure goes to 1 atmosphere, and the temperature drops back to −78 °C. Keeping the bulb sealed permits the pressure to rise until the bulb breaks. Room temperature is around the critical temperature of carbon dioxide, 31 °C. A tank of carbon dioxide at room temperature is filled with gas or a mixture of gas and liquid.

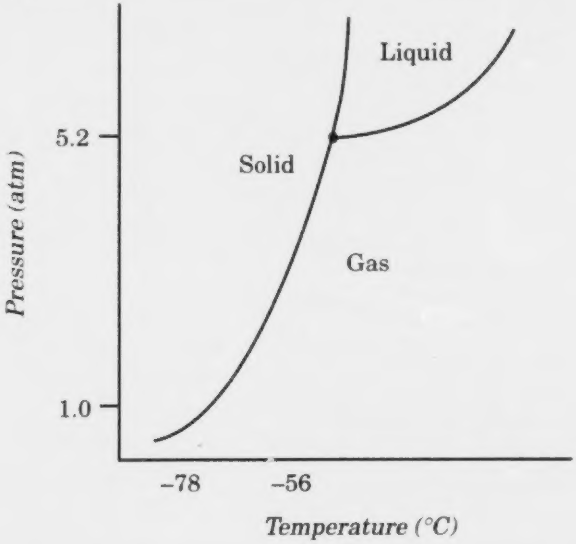


Figure 16. Schematic phase diagram for CO_2.

Reference

Adapted from Becker, R. (1991). Wet dry ice. *Journal of Chemical Education, 68*, 782-783.

Metaphors and Analogies

The three phases of matter can be likened to people attending a football or basketball game. When the fans are all seated before the game or standing for the national anthem, they are arranged like the particles in the solid phase of matter. At half-time when the fans get up and mill around in "clumps" to go to the refreshment stand or the bathrooms, they are behaving like liquid particles. When the game is completed, the fans leave the sports arena, get in their cars, and head randomly in all directions. They are scattered over a very large area compared to the area which they occupied at the game. In this case, they are behaving like gas particles. Note: If you find that the pipet ruptures before the triple point is reached, try Tygon™ tubing, clamping a 6-in. piece at both ends.

Pictures in the Mind

1. Draw a picture of the solid and liquid particles in a container of a substance at its melting point. *[The important differences are in the rigidly arranged structures of the crystalline solids and the more loosely arranged liquid particles. At a given temperature, there is no difference in the average kinetic energy of the solid or liquid phase. The major difference at the melting point is that liquid particles possess mainly translational kinetic energy while a major portion of the kinetic energy of the solid particles is vibrational in nature.]*
2. Draw pictures of solid particles, liquid particles, and gaseous particles for a particular substance. *[The student answers should show that they understand the differences among the three phases of matter in bonding, order, and amount of empty space.]*
3. Draw a graph of temperature *vs.* heat input (a warming curve) for water under normal atmospheric pressure. Take the water from solid ice at –20°C to gaseous steam at 120 °C. [*The student graphs should indicate they understand that there are three "slanted" portions to the graph (kinetic energy changes) and two "flat" portions to the graph (potential energy changes). If they are more sophisticated in their understanding of this concept, they may also indicate differences in heat capacities of ice, water, and water vapor by controlling the slope of the "slanted" portions. The differing heats of fusion and vaporization could be indicated by showing the "flat" portions of the curve at the phase changes to be different lengths.]*

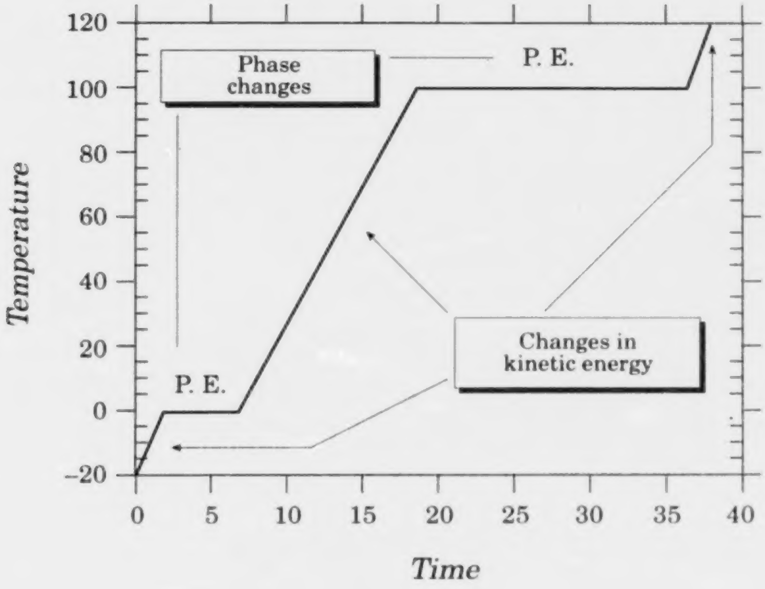


Figure 17. Warming curve for water.

Language of Chemistry

Glossary

Archimedes' principle weight of an object submerged (completely or partially) in a fluid is lowered by the weight of the volume of fluid it displaces.

atmospheric pressure pressure exerted by the atmosphere. Though a function of weather conditions, standard pressure is defined as the pressure that will sustain a 760-mm column of mercury metal in a barometer. The SI equivalent of one atmosphere is 101.3 kilopascals.

attractive forces pulls between objects causing them to resist separation from one another; usually used when discussing chemical bonds and attractions between atoms, ions, and/or molecules.

boiling vigorous change from liquid to gaseous state characterized by formation of bubbles of vapor beneath the surface of the liquid.

boiling point temperature at which boiling occurs.

calorie unit of energy. One calorie is 4.18400 joules; this is about the amount of heat required to raise the temperature of 1 g of water at 15 °C by 1 °C.

capillary rise rise of a column of fluid in a tube, usually a cylindrical tube of small dimensions, due to the surface tension of the fluid.

closed system system in which there is no flow of energy or matter into or out from the systems, such as a covered, sealed, insulated Thermos™ bottle.

condensation process of going from gas state to a condensed state, usually the liquid state.

contraction shrinkage of the dimensions of a solid or liquid due to temperature change, usually cooling.

cooling curve curve plotting temperature as a function of either energy removed or time.

critical temperature temperature above which only the gaseous state of a substance may exist regardless of the amount of pressure applied.

crystal lattice imaginary three dimensional array of points used to describe the positions of atoms in a crystal.

density ratio of the mass of a system to its volume; an intensive property.

dynamic equilibrium situation characterized by changes in particular particles but with no net change in the total number of particles in a given state. Example might be a bank in which the total number of 20 dollar bills remains the same between open and close, and remains nearly constant throughout the banking day, but has bills with different serial numbers at the end of the day as compared with the beginning as the result of bank transactions.

expansion increase of the dimensions of a solid or liquid due to temperature change, usually by heating.

extensive properties properties that depend upon the amount of material measured, such as mass, volume, heat, area, and moles.

freezing change from the liquid to the solid state.

heat capacity the ability of a system to hold heat energy.

heat of fusion the amount of energy required to bring about the melting of a solid at its normal melting (freezing) temperature.

TIPS FOR THE TEACHER

heat of vaporization the amount of energy required to bring about the boiling of a liquid at its normal boiling (condensation) temperature.

hydrogen bonding attractive force resulting from a hydrogen atom bridging between two small, electronegative atoms (such as F, O, or N atoms).

ice solid water.

intensive property a property that does not depend upon the amount of sample present, such as density, concentration, pressure, or specific heat.

intermolecular forces forces between molecules holding the molecules together in an aggregate; responsible for formation of condensed states in molecular substances.

kilojoule 1000 joules; one joule is the kinetic energy of a 1 kilogram object moving at a speed of 1 meter per second.

kinetic energy energy of motion.

latent heat heat of fusion or heat of vaporization; heat released (or taken in) during a change of state.

LeChatelier's principle when a stress is placed upon a system at equilibrium, the system shifts its equilibrium position so as to relieve the stress.

mass a fundamental property of matter that makes it subject to gravitational force. It is measured as a body's tendency to resist changes in its state of motion.

melting change from the solid to the liquid state.

melting point temperature at which melting occurs.

non-Newtonian fluid a fluid whose viscous behavior is not typical of common liquids but whose viscosity increases with increase in rate of flow.

phase change change from one phase to another. When applied to substances not undergoing chemical changes, this refers to melting, freezing, boiling, condensing, or subliming.

phase diagram plot of pressure versus temperature for a substance indicating the states of matter present at each point by labelling regions as solid, liquid, or vapor. Several solids and sometimes more than one liquid are possible.

physical change change that does not involve changing chemical bonds such as change of state for a substance.

physical property property relating to physical changes or state (such as melting point or density).

potential energy energy of position. Latent heats are examples of potential energy, as are changes in energy due to changes in chemical bonding.

pure substance one element or compound.

slope the ratio of the increase in the y-increment for a given x-increment in a straight line graph.

specific gravity the ratio of the mass of a given volume of material to the mass of the same volume of water.

specific heat the amount of heat required to raise the temperature of one gram of a material by one Celsius degree.

sublimation change from the solid to the gaseous state.

surface tension molecules on the surface of a liquid experience different forces than those in the interior that are surrounded by other molecules. The surface acts as if it experiences a tension or force as a result.

temperature change difference between final and initial temperature.

triple point temperature and pressure at which solid, liquid, and gaseous states of a substance coexist and remain at equilibrium with one another. A unique point on a phase diagram.

vaporization process of going from liquid (and sometimes solid) to gaseous state.

vapor pressure pressure exerted by vapor molecules in equilibrium with corresponding liquid or solid.

viscosity resistance of a liquid to flow.

volatile liquid liquid that readily evaporates. Indicates weak intermolecular forces.

volume the amount of space occupied by a system. The amount of room it takes up.

warming curve curve plotting temperature as a function of either energy added or time.

Vapor Pressure

Lubricating greases have very low vapor pressures due mainly to the high molecular masses of their molecules. Consider how long one would expect to wait to see a container of lubricating grease completely "vaporize" while resting open in the family garage.

But in outer space, the atmospheric pressure is so low that the vapor pressure of normal lubricating grease is higher than this low atmospheric pressure. Thus, normal lubricants rapidly boil away and the ensuing friction causes machinery to fail. Needless to say, special lubricants have been developed for use in our space program.

Clothing can be dried on the clothesline even in the winter when the ambient temperature is below freezing because water has vapor pressure even when solid (4.6 mm Hg at 0°C). At temperatures below freezing, the moisture in the clothing sublimes away.

The fact that water has substantial vapor pressure in the solid state explains why one must wrap steaks in a "zip lock" bag or other tight wrap to preserve the water content. Failure to protect the steak may lead to freezer burn or a dry "beef jerky" steak.

Under prevailing atmospheric pressure, dry ice (solid CO_2) sublimes to the gaseous state. The triple point pressure of carbon dioxide is above atmospheric pressure. Under normal conditions, liquid carbon dioxide will never be seen.

Fruit juice concentrates which can be purchased frozen at grocery stores are concentrated by boiling the major portion of the water out of the "fresh squeezed" juice. If the boiling were done at normal atmospheric pressures, the high temperature would destroy molecules which provide the fresh flavor. Vats of juice are placed under a partial vacuum and boiling takes place at a lower temperature. The molecules are not denatured but are preserved to provide taste at the breakfast table.

Everyone has heard of freeze dried coffee. Interestingly, this same technique can be applied in taxidermy. The animal specimen to be preserved is first carefully frozen. Then the specimen is placed in a vacuum chamber and the atmospheric pressure is lowered until the ice in the tissue sublimes away. Shrinkage and distortion of the specimen, which occur under normal drying techniques, are reduced to a minimum.

The white granulated sugar purchased at your local grocery store is produced by heating sugar syrup solution under a vacuum until it becomes supersaturated and deposits crystals around "seed" crystals. Heating under a vacuum reduces the operating temperature and minimizes discoloration (caramelization) due to decomposition of the sugary syrup.

Boiling Point

Liquids will not boil until their vapor pressures have increased to the prevailing atmospheric pressure. Pressure cookers take advantage of this fact. The closed pressure cooker causes the pressure to increase substantially within the vessel. Vapor pressure is directly related to the temperature of the liquid. This pressurized water in the pressure cooker must be raised to a higher temperature before it begins to boil, hence the food is being subjected to a higher temperature and will cook faster than in an open container.

Every camper who loves the mountains and prefers a "three minute" egg in the morning has faced the problem of cooking under reduced atmospheric pressure. Since the atmospheric pressure is lower at higher altitudes, water will boil under a lower pressure at the mountainside campsite. An egg boiled for three minutes under these conditions will be "runnier" than the camper is accustomed to at home.

Cake bakers in the mountains also realize that an extra quantity of water must be placed in the cake mix to make proper batter. Since the water boils at a lower temperature under these lower atmospheric pressure conditions, more water boils away during the baking process. If the normal recipe is followed, the cake will not have the desired moisture level.

Latent Heat

When one gram of water freezes, 334 joules of heat are released with no change in temperature. This heat is not apparent because it is released with no temperature change; it is called latent heat. Latent means hidden. Latent heat of fusion is actually potential energy released as the liquid water molecules bond together to form solid ice.

Fruit and vegetable growers use this reservoir of latent heat to protect their crops when freezing weather threatens. As water sprayed on the crops cools and freezes, the latent heat released in the cooling and freezing process is released to the atmosphere around the crops. This supply of heat keeps the ambient temperature higher than it would have been if water were not sprayed on the crops.

For every gram of water which is evaporated from the surface of the earth by solar heating and converted into water drops in a rain cloud, a minimum of nearly 2500 joules of kinetic and potential energy are released into the atmosphere. This tremendous amount of energy produces such damaging phenomena as high winds, lightning, hail, tornadoes, and hurricanes.

Heat Capacity

Water has an extremely high heat capacity (highest of any substance). Much energy is required to speed up water molecules and in turn, these molecules release much energy as they slow down and cool. This is why a hot water bottle is such a comfort on a cold winter's night.

Because of the high heat capacity of water (warms slowly—cools slowly) areas near oceans or seas have moderate climates—no really hot summers or frigid winters. A case in point is the southern coast of Ireland which is bathed by the Gulf Stream. This high heat capacity water carries heat all the way from tropical areas to Ireland. Southern Ireland, although it is at a latitude above that of the U.S/Canadian border, has year-round temperatures high enough to support subtropical growth including palm trees.

The exaggerated temperatures experienced in desert areas of the earth are due in great part to the lack of moisture (water) in the soil. With little water in the sand, rather than being stored in the soil, heat from the sun radiates rapidly back into the atmosphere to heat it. At nightfall, the soil has stored little heat to radiate to the atmosphere to moderate nighttime temperatures.

The heat capacity of aluminum is approximately 1/5 that of water (aluminum warms faster and cools faster than water). Consider that although the aluminum pan containing a TV dinner may be cool enough to the touch to be carried to your easy chair, the first bite of the enclosed food items has a tendency to burn your tongue. Blame it on the high water content. There is a lot of heat released as these hot water molecules cool in your mouth.

Expansion/Contraction

In general, materials expand when heated because of the increased kinetic energy of their particles at higher temperatures. Higher kinetic energy results in higher velocities ($E_k = 1/2mv^2$, where E_k = kinetic energy, m = mass and v = velocity). When the particles strike one another at higher velocities they bounce further apart. This fact accounts for the expansion exhibited by most materials when heated. Likewise, the decrease in kinetic energy and correspondingly lower velocities account for the contraction of most materials as they cool.

Temperature measuring devices (mercury, alcohol, gas, and bimetallic strip thermometers) rely upon the regular expansion and contraction of materials during heating and cooling.

Aircraft once were hot riveted because hot metal is more malleable. As hot rivets cool, they contract and fit loosely in the holes in the metal frame. This fact was not a problem until aircraft became supersonic. During the buffeting of passage through the sound barrier, the loose fitting hot rivets led to vibrations and cracking of the airframe. Several aircraft and pilots were lost as a result of this "slack" problem. Supersonic aircraft are now assembled through a cold riveting process. As these cold rivets warm to normal flight temperatures, they actually expand to fit tightly into the holes in the metal frame.

Liquids when heated expand at faster rates than do solids. If one fills an automobile gasoline tank (solid) to the brim with gasoline (liquid) on a cool summer evening, and doesn't drive the auto until the hot part of the next day, gasoline will be flowing out of the filler tube and onto the driveway. The liquid, gasoline, and the solid, gas tank, have both expanded, but the liquid has expanded more. Thus an amount of liquid flows out of the warmed tank.

The fact that materials (including glass) expand and contract due to temperature changes is the reason that volumetric flasks are accurate only at one designated temperature.

Pyrex glass can withstand rapid temperature changes better than normal (crown) glass because it expands and contracts at about 1/3 the rate of normal glass. If one pours cold liquid into a hot crown glass container, the inner surface cools and contracts rapidly relative to the still hot outer surface. The stress is normally too much for the forces within the glass to withstand, and it cracks. With Pyrex® glass, given the same temperature differential between the inner and outer surfaces, the contraction rate is only 1/3 as great. This lower stress is within the limits which the bonding structure can tolerate, and the Pyrex® container resists this cold liquid shock.

Hydrogen Bonding in Water

The hydrogen bonding structure in water produces a very open hexagonal crystal lattice which requires that water molecules move farther apart in order to solidify into ice. Thus, water goes against the normal trend and expands as it freezes. As a result of this unusual behavior, ice is less dense than water and floats on water. If not for this uncommon property, ice would form on the bottom of lakes and oceans in the higher latitudes of the earth. Freezing would occur from bottom to top and aquatic life would find it impossible to survive.

Another material which exhibits this unusual expansion when freezing is type metal. If type metal contracted as it became solid, it would pull away from the type mold, and "chunks" of the letters' surfaces would be missing. The resulting print would be fuzzy and difficult to read. Type metal expands upon freezing to fill every nook and cranny in the type mold.

Normally, substances contract when they enter the solid state. Placing enough pressure on such substances in the liquid state will cause them to become solid. Water, which expands upon freezing, exhibits reverse characteristics. Solid water (ice) when pressurized will return to the liquid state. This is why those icy footprints appear on your sidewalks if pedestrians get to them on a winter morning before you can scoop the snow. This phenomenon leads to the glacial "flow" that is characteristic of large bodies of ice.

The open structure of the crystal lattice of ice is both a blessing and a curse. The expansion of ice is an indispensable aid in the weathering process. Freezing water expands in small fissures in rocky surfaces and these cracks grow larger. After a fashion rocks are reduced to the soil necessary to sustain life. That's small comfort to the home owner who must pay several thousands of dollars to replace sidewalks and driveways which have been converted to rubble by this same weathering process.

Pattern Recognition

1. All "slanted" lines on warming or cooling curves indicate changes in kinetic energy (changes in average velocity of the particles). The slopes of these lines indicate the magnitude of the specific heat (the greater the slope, the lower the specific heat).
2. All horizontal lines on warming or cooling curves indicate changes in potential energy (significant changes in attractive forces). The length of these lines reflects the strength of the attractive forces; long horizontal portions suggest strong attractive forces.
3. All portions of a warming curve indicate an increase in energy for the substance being heated. As time goes on during the warming process, energy is being stored in the substance.
4. All portions of a cooling curve indicate a decrease in energy for the substance being cooled; energy is released by the substance as it cools.
5. Substances with low vapor pressures have strong attractive forces in the liquid phase. These substances will have high boiling points and will also evaporate slowly.
6. Substances with high vapor pressures have weak attractive forces in the liquid phase. These substances will have low boiling points and will evaporate rapidly.

Common Student Misconceptions

1. **"At the melting point of a crystalline substance, the particles that have turned liquid are moving faster than the particles that are still in the solid phase."**

   In fact, the difference between the solid and liquid particles at the melting point is a difference in the forms of kinetic energy (translation, rotation, vibration). The liquid particles also have added potential energy due to the fact that they have been moved a distance apart.

2. **"The phase in which substances exist at normal temperatures is the phase in which they always exist."**

   For example, students find it difficult to picture iron in the gaseous state or mercury in the solid state. They usually think of water only as a liquid, and need to be jogged to think of ice or water vapor. At liquid nitrogen temperatures, it is possible to drive a rubber nail into a wooden board with a mercury hammer.

3. **"Materials which are more viscous are necessarily more dense."**

   For example, ask students which is more dense, milk or cream. They will frequently conclude that cream is more dense because it is "thicker." Yet, cream floats on milk and thus is less dense. They may also draw the same faulty conclusion when comparing oil and water or vinegar and oil. We can think of viscosity as a rate process where increasing energy (temperature) allows molecules to move past their neighbors more rapidly. Viscosity is partially determined by intermolecular forces of attraction, and thus parallels energies of vaporization.

4. **"The freezing point and boiling point of substances are fixed points and never vary. (Water always boils at 100 °C.)"**

   The freezing and boiling points are dependent upon many things, especially the prevailing pressure and the purity of the substance.

5. **"Substances are always cold when they 'freeze.'"**

   Students never seem to think of iron "freezing" at a very high temperature, but that is exactly what any material does as it passes from the liquid to solid phase, no matter what the temperature is when this process takes place.

6. **"Ice never gets colder than 0 °C."**

   This is a very common idea held by students. To deal with this misconception, place water in a test-tube. Place a thermometer in the water. Place the whole system in a beaker with an ice/salt/water mixture. Have the students watch the temperature. This would also work nicely as a computer/temperature probe demonstration

Decision Making

Place one or two ice cubes in a beaker of water. Carefully pour water into the beaker until the water is "heaped up" due to the surface tension of the water. Ask the students whether or not the water will overflow and run onto the table top when the ice cubes melt. As the ice cubes melt, lead a discussion of "What?" the students think will happen and "Why?" The answer will become self-evident as the cubes melt.

Assessment

1. **List the two condensed states of matter and describe each state.**

   The two condensed states of matter are the solid state and the liquid state. Solids and liquids, unlike gases, are not compressible since the particles in solids and liquids are drawn closely together by quite strong, short range attractive forces.

   The solid state is that phase of matter with the lowest total energy. Solid particles are arranged in a regular pattern called a crystal lattice, which allows the precise prediction of the position of nearly all other particles in the solid by knowing the positions of only a few particles. Because of the strong attractive forces in solids, solids have both definite shapes and definite volumes.

When enough energy is added to the solid state, the very strong attractive forces are overcome and the crystal lattice is disrupted, causing the material to melt. The liquid state is the result of this process. In the liquid state, the particles possess more potential energy and more rotational and translational kinetic energy. The particles in a liquid are thought to be in close contact with one another, with nearest neighbors arranged around one another in fairly regular ways. The particles can be envisioned as moving around in "clumps." The arrangement is certainly not as regular as in a solid and the attractive forces in a liquid are not as strong. This accounts for the fact that liquids do not have a definite shape.

2. **Four samples of the same material were measured. Use the following data (Figure 18) to determine the density of a material.**

Sample No.	1	2	3	4
Mass of object (in air) (g)	17.78	4.08	27.07	5.40
Mass of object (in water) (g)	11.31	2.56	17.04	3.38

Figure 18. Sample data for density determination.

According to Archimedes' principle, a body immersed in a fluid loses an amount of weight equal to the weight of the fluid displaced. Water has a density of 1.00 g / cm^3. Thus, every gram of reduction in the weight due to this Archimedes' buoyancy effect represents a volume of 1.00 cm^3 for the immersed object. Below is an extended data chart revealing the densities of the four test objects.

Sample No.	*1*	*2*	*3*	*4*
Mass of object (in air) (g)	*17.78*	*4.08*	*27.07*	*5.40*
Mass of object (in water) (g)	*11.31*	*2.56*	*17.04*	*3.38*
Difference in mass (g)	*6.47*	*1.52*	*10.03*	*2.02*
Volume of sample (cm^3)	*6.47*	*1.52*	*10.03*	*2.02*
Density of sample (g / cm^3)	*2.75*	*2.68*	*2.70*	*2.67*

Figure 19. Density of objects.

NOTE: The density of the samples was calculated by dividing the mass of the object in air by the volume of the sample as determined by water displacement. *For example, since D = m / V, the density of the first sample would be D = 17.78 g / 6.47 cm^3 or 2.75 g / cm^3. The average density of these four test samples is 2.70 g / cm^3. These data represent trials conducted on aluminum metal. A valuable addition to this test question would be to ask the students to determine the average density, and to predict what the unknown metal is by using a table of densities from a textbook or suitable handbook.*

3. **Describe the physical and energy changes that take place as water vapor condenses and then freezes.**

   The particles in water vapor (the gaseous state) are in random motion. As heat energy is removed from the water vapor molecules, their velocities decrease and they collide less violently. At a certain point, the energies in the collisions are reduced enough that the attractive forces exerted on one another by the vapor molecules cause those molecules begin to "clump" together. This is the condensation point. During the condensation process, the vapor molecules move closer together and a great deal of potential energy is released. This potential energy is known a latent heat of vaporization. During the condensation process, the kinetic energy of the system is not reduced, thus the temperature remains constant. The flat portion of the characteristic cooling curve for water is formed during this process.

   Once all of the water vapor has condensed into liquid water, the molecules continue to lose heat energy as the clumps of molecules slow down, thus decreasing the amount of translational and rotational kinetic energy. Since the average kinetic energy is decreased during this process, the temperature will fall. The "slanted" portion of the characteristic cooling curve for water is formed during this process.

   When the "clumps" of water molecules slow sufficiently, the attractive forces cause the water molecules to be arranged in a regular crystalline lattice. The average kinetic energy will not decrease. The latent heat of fusion released at this point comes from a decrease in potential energy representing the change of the water molecules from the liquid state to the regular arrangement in the ice crystal lattice. The second flat portion of the characteristic cooling curve for water is formed during this energy release.

   Once the water is a solid, it has very little potential energy remaining. It is in a very regular crystal structure. The major portion of its remaining energy is in the form of kinetic energy which the water would lose as it is cooled toward absolute zero. The final slanted portion of the characteristic cooling curve for water would be formed during this energy release.

4. **5-mL water at 25 °C are injected into an evacuated, sealed, insulated 1-L container. Describe the changes that are expected. Provide suitable explanations for these changes.**

   A portion of the water evaporates until the pressure of water vapor inside the container reaches the equilibrium vapor pressure. Once this pressure is reached, a dynamic equilibrium is maintained. Some water molecules evaporate while others condense; the total amount in the vapor state remains constant, and the total amount in the liquid state remains constant.

5. **Provide a semiquantitative prediction of the boiling temperature of water at a pressure of 1.05 atmospheres. Explain your prediction.**

   At 1.05 atmospheres, the boiling temperature of water will be higher than 100 °C. As the external pressure is increased, the temperature at which boiling takes place (vigorous evaporation throughout the body of the liquid) is increased.

6. **Predict whether water or methanol has a higher vapor pressure at 25 °C. Provide an explanation for your prediction. Check your prediction using a handbook. If your prediction was incorrect, try to see if you can provide an alternative explanation for the effect. Explain what might have been incorrect with your first explanation.**

Methanol has a lower normal boiling point than water and is expected to have a higher vapor pressure at room temperature. From a phenomenological perspective, methanol is closer to boiling at room temperature than is water, so its vapor pressure is higher. The underlying explanation is usually given in terms of the more extensive hydrogen bonding available to water as compared to methanol, thus making evaporation (overcoming attractive forces) from methanol an inherently easier process than evaporation from water.

HISTORY: ON THE HUMAN SIDE

1. **Lavoisier and the Ice Calorimeter**. Antoine Lavoisier and French mathematician LaPlace designed an apparatus called an ice calorimeter to measure heat released in various combustion reactions including animal respiration processes. Animals were placed in the calorimeter and the amount of ice melted was compared to the amount melted as charcoal was burned in a similar calorimeter. Oxygen intake and carbon dioxide (fixed air) production were also recorded and compared to the ratios of these two gases involved when the charcoal burned. Lavoisier and LaPlace concluded that the combustion of food in the animal's body supplied the heat to maintain its body temperature much as burning charcoal released heat through its oxidation.

2. **Joseph Black and Latent Heat**. In the 18th century it was commonly believed that when ice or snow reached its melting point, only a small addition of heat was necessary to convert the solid into liquid water. Joseph Black, a Scottish chemist, conducted a series of experiments which indicated that once ice began to melt, the temperature of the ice/water mixture remained constant so long as any ice remained. He calculated that the amount of heat necessary to melt ice was about 80% of the heat necessary to raise that same water from the freezing point to the boiling point—by all measures not a small amount of heat. Black coined the term "latent heat" (hidden heat) to refer to this heat which entered the ice/water mixture without raising the temperature.

   Black went on to show that absorption of a substantially larger amount of heat occurred in a similar fashion as water was converted to water vapor. These findings were subsequently applied by fellow countryman James Watt as he improved Thomas Newcomen's steam engine by designing a separate chamber in which to condense the waste steam from the steam engine.

HUMOR: ON THE FUN SIDE

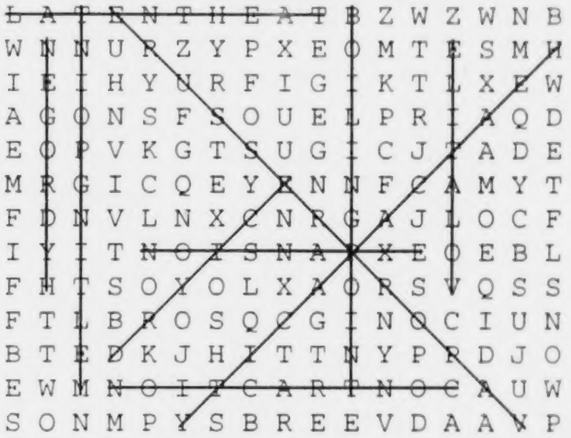
1. Rhode Island is a condensed state!
2. What is H_2O_4? [What is H_2O for?]

   [Answer: What ever grabs you. For bathing...for watering plants...for fish to live in...etc.]
3. What is the solid state?

   [New Hampshire, the Granite State!]
4. **A thought**: Does SnO = $H_2O(s)$?

   CHEM 13 NEWS, December 1984, p. 824

5. Word Search (see *Appendix* for master copy)



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L A T E N T H E A T B Z W Z W N B
W N N U R Z Y P X E O M T E S M H
I E I H Y U R F I G I K T L X E W
A G O N S F S O U E L P R I A Q D
E O P V K G T S U G I C J T A D E
M R G I C Q E Y E N N F C A M Y T
F D N V L N X C N R G A J L O C F
I Y I T N O I S N A P X E O E B L
F H T S O Y O L X A O R S V Q S S
F T L B R O S Q C G I N O C I U N
B T E D K J H I T T N Y P R D J O
E W M N O I T C A R T N O C A U W
S O N M P Y S B R E E V D A A V P
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Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. This quantity is exerted by a gas that is in equilibrium with its liquid state at a given temperature (2 words).
2. Temperature at which a liquid's vapor pressure equals the atmospheric pressure (2 words).
3. Energy that is released but not apparent since it goes toward changing phase (2 words).
4. The quantity of energy that a given object can absorb per degree Celsius change in temperature (2 words).
5. What generally occurs when solid materials are heated due to increased kinetic energy of the particles.
6. Volume change that normally takes place when materials are cooled due to lower kinetic energies of the particles.
7. Type of bond that gives water its unusually high boiling point.
8. Adjective describing liquid that has a high vapor pressure and readily evaporates at room temperature.
9. The temperature at which a solid changes to its liquid form (2 words).
10. Common name for $CO_2(s)$ (2 words).

Answers: 1. VAPOR PRESSURE 2. BOILING POINT 3. LATENT HEAT 4. HEAT CAPACITY 5. EXPANSION 6. CONTRACTION 7. HYDROGEN 8. VOLATILE 9. MELTING POINT 10. DRY ICE

6. See cartoons at end of module.

MEDIA

1. *Crystals and Their Structures,* CHEM Study film/video available from Ward's Natural Science Establishment, Inc., 5100 West Henrietta Road, P.O. Box 92912, Rochester, NY 14692-9012; (800) 2660.
2. The *World of Chemistry* videotapes secondary school version is available from WINGS for Learning/SUNBURST, 101 Castleton Street, Pleasantville, NY 10570; (914) 747-3310; (800) 321-7511; (914) 747-4109 (FAX). "Module 3: States of Matter" and Module 1: Water" are appropriate for this module.

3. Software published by *JCE:* Software, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   a. *Bravais*, by Philip Pavlik. Vol. III B, No. 1, for IBM PS/2 PC-compatible computers.

   b. *Animated Demonstrations*, by Philip Pavlik. Vol. V B, No. 2, for Windows/ IBM PC.

   c. *Crystal Lattice*, by David Trapp. Vol. IV A, No. 1, for the Apple IIGS computer.

4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   For the IBM PS/2 PC-compatible computer: PC 2803

5. Videodisc published by *JCE: Software*, a publication of the *Journal of Chemical Education,* Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Liquid Nitrogen and Racquet Balls," "Changes of State in Bromine," and "Water and Ice," three chapters on *The World of Chemistry: Selected Demonstrations and Animations: Disc I* (double sided, 60 min.). Special Issue 3.

6. Written materials and devices distributed by ICE–Institute for Chemical Education, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-3033 (voice) or (608) 262-0381 (FAX).

   a. The Solid State Model Kit

   b. The Optical Transform Kit

Links/Connections

WITHIN CHEMISTRY

This particular module would fit well with units concerning gas laws, crystal structures, and thermochemistry. The concept of density often is included within introductory units. The laboratory work lends itself to data collecting and interpreting with a computer.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

Physicists are very interested in latent heat of fusion, specific heat, density, and Archimedes' principle. Technology has a vested interest in the solid state and as such represents a specialized field of study to physicists. Water is a topic in itself that is of special interest to biologists.

TO THE CONTEMPORARY WORLD

New and improved building materials may lead to future industrial changes. Better understanding of condensed states has provided liquid crystals and may be the key to improvements on superconductivity.

(See *Language of Chemistry* for additional examples of applications in the contemporary world.)

Community Resources

1. **Field trips**. Local Weather Service, water treatment plant, steel mill or foundry, food processing plant (note Societal Issue #1), museums (gem and mineral collections).
2. **Knowledgeable Individuals**. Meteorologists, metallurgists, mineralogists, materials engineers, hydrologists, jewelers and gem hobbyists, X-ray crystallographers, civil engineers, mechanical engineers.

Societal (STS Issues)

1. According to University of California Cooperative Extension Service statistics, it takes 4533 gallons of water to produce the food a typical American eats in a day. The water use total includes all water necessary from growth of the food product to processing to cooking and placing the finished product on your dinner plate.

   Listed are a selection of food products indicating the *number of gallons of water needed to produce one pound* of each item for the table.

Food	Gallons of H_2O	Food	Gallons of H_2O
Apples	49	Grapes	70
Almonds	188	Onions	16
Beef	5214	Oranges	65
Cantaloupe	51	Pork	630
Celery	22	Potatoes	24
Chicken	815	Rice	203
Corn	122	White Bread	122
Dry Beans	326		

According to the *Examiner* article, "The consumer could cut the water use considerably if he got rid of the beef, about four bites of which use enough water to fill the average hot tub, and substituted, say a soy burger."

Discuss with your class the *Examiner's* statement in light of the increasing pressures upon water resources in this country.

Reference: *San Francisco Examiner*. (1990, June 24). p. B-1.

2. Polystyrene plastic in the form of Styrofoam® cups and containers, is less than 1% of the solid waste stream in the United States. In many locations in the country, there is a rush to replace Styrofoam® packaging with paper because of the concern that plastics aren't biodegradable.

   Most landfill garbage, including paper, undergoes little biodegradation. A landfill which is environmentally sound does not allow the air and light necessary for biodegradation. The only paper which properly biodegrades is that "littered" along roadsides where air and sunlight are abundantly available. Since neither the paper nor plastic degrade in a typical landfill, and since paper is 40 times more abundant in landfills, maybe Styrofoam® plastic is not as big a culprit as most people believe. A better solution to the waste problem (paper or plastic) may be recycling. Discuss the pros and cons of this statement with your chemistry students.

   This might be a good place to do some extra library research before coming to a decision about this problem. Divide the class up into pro/con Styrofoam® groups, study it at the library, and have a lively class discussion.

Extensions

An excellent Piagetian laboratory exercise, *Some Like It Hot!*, is provided as an extension to the heat of fusion laboratory. It is highly recommended as a way to broaden students' concepts of heat and heat flow.

In addition, you may wish to have interested students look into the ten topics listed below for a broader perspective on the condensed states of matter.

1. Anhydrous substances
2. Desiccants
3. Hygroscopic, deliquescent, and efflorescent substances
4. Diamond and graphite structures (allotropes)
5. Liquid crystals
6. Semiconductors
7. Hydrogen bonding in water and related properties of water
8. Surface tension and capillarity
9. Phase diagrams
10. Liquefaction of gases

LABORATORY ACTIVITY: STUDENT VERSION

Extension Activity: Some Like It Hot!

Introduction

When samples of water at different temperatures are mixed, the final temperature of the mixture will have a value that lies between that of the hotter sample and the colder sample. According to the laws of thermodynamics, heat flows from warmer into cooler bodies. The warm water cools down and the cool water warms up. The process continues until a uniform temperature is reached at which time the temperature change ceases.

Purpose

Your task is to find the quantitative relationship that relates the final temperature of the mixture to the initial temperatures of the cold and hot samples, respectively.

Safety

1. Wear protective goggles and apron throughout this laboratory activity.
2. As always when working with hot water, be careful not to come in contact with the hot water or the hot containers in which it is held. Move all hot objects with beaker tongs or a hot pad.

Procedure

1. Use the following procedure to mix measured volumes of hot and cold water. As a guide, be sure that the final volume of the mixture is between 100 mL and 150 mL.
2. Place a measured volume of cold water, as determined by using a graduated cylinder, into one Styrofoam® cup.
3. Place a measured volume of hot water into the other Styrofoam® cup.
4. Measure the temperatures of the cold water, and the hot water, to the nearest 0.2 °C.
5. Pour the contents of one cup into the other cup, and measure the temperature of the mixture.
6. Record the data into a table similar to the one below.

Trial	Volume (Cold)	Temp (Cold)	Volume (Hot)	Temp (Hot)	Volume (Mix)	Temp (Mix)
1						
2						
3						

7. Repeat the experiment two more times using different volumes of hot and cold water.
8. Use the data to help you formulate a rule that will allow you to predict the temperature of the final mixture. If you have a rule, share this with your instructor, who will ask you to predict the temperature of a mixture you have not investigated.

9. Whether or not you have established a valid rule, the second task of this experiment is to use the data from one of your three trials to construct graphs. From the selected trial, make three graphs using the temperature of the sample versus the volume of the sample for the cold, hot, and hot and cold mixture, respectively. To make the graphs, plot a single point (*e.g.*, Volume Cold, Temperature Cold) on the graph paper. Construct a rectangular area by drawing the respective lines going to the x and y axis. Be sure to use the same increments on the axis scale for each graph.

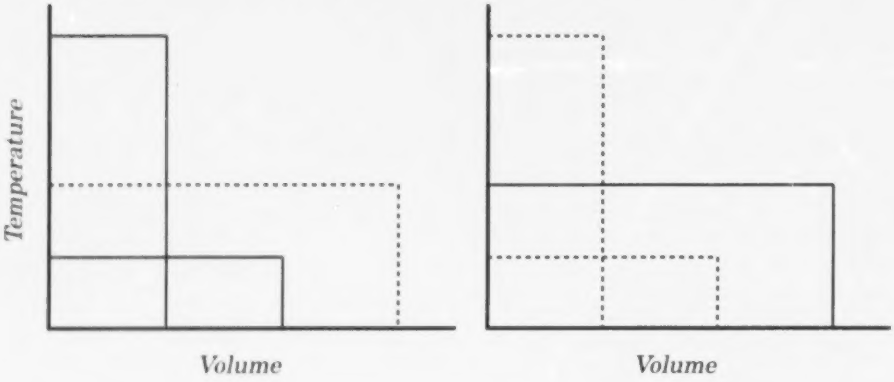

Figure 20. Sample for 20 mL of 40 °C water mixed with 40 mL of 10 °C water.

10. Using the axis readings determine the area of each rectangle. Find a relationship between the sum of the area associated with the hot and cold water and the area associated with the mixture.
11. Answer Questions 1 to 3 from the *Data Analysis and Concept Development* section based upon your data and observations so far.
12. The third task is to put your rule to a laboratory test. First prepare a thermometer by partially covering it with black paint or paper or opaque rubber tubing. Carefully slide the tubing over the top portion of the thermometer, leaving the lower portion of the thermometer exposed up to the 40 °C range. (This will prevent the student from measuring directly the temperature of the water in the coffee urn.) There will be a source of hot water in the laboratory in a coffee urn or insulated cooler. You are to determine the temperature of this hot water in the container by indirect means. Take some time to plan the procedure of this aspect of the experiment. Plan to do at least three trials.
13. Finally, check with your instructor and compare the actual temperature with your prediction. Discuss how the experimental design could be improved to increase accuracy. You may test your new design by repeating the experiment.
14. Answer Questions 4 to 7 in the *Data Analysis and Concept Development* section.
15. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. State the relationship for the three graphs.
2. Write the relationship as an algebraic formula.
3. Suppose you were given 205 mL of 30.0°C water that is mixed with 101 mL of 75.0 °C water, what is the final temperature of the mixture?

4. Assuming that the temperature of water in the coffee urn or insulated cooler didn't change, find the per cent error.
5. How good was your precision and accuracy in your originally designed experiment?
6. How could or did you design the procedure to improve accuracy?
7. Suppose you mixed 105 mL of 15.0 °C water with 70.0 mL of unknown temperature, and the final temperature of the mixture reached 43.0 °C. What was the initial temperature of the hotter water?

Reference

This activity was originally developed at SUNY Cortland; this version was enhanced by John Ricketts of DePauw University.

Implications and Applications

Mixing liquids of different temperature and predicting the final temperature is an everyday phenomenon. We do it when we add milk to our coffee. Whenever there is a power plant and also in many factories, waste heat is removed through cooling water that is dumped into the environment, often into a passing stream or river. This leads to a temperature increase, often one leading to serious ecological consequences.

Heat is exchanged between two large bodies of different temperature that are brought into contact. In countries where malnutrition abounds, breast milk is often the only source of energy for a baby. However, the mother also provides heat while nursing (and at other times) that reduces the baby's total energy needs.

LABORATORY ACTIVITY: TEACHER NOTES

Extension Activity: Some Like It Hot!

Major Chemical Concept

When samples of water at different temperatures are mixed, the final temperature of the mixture will have a value that lies between that of the hotter sample and the colder sample. According to the laws of thermodynamics, heat flows from warmer into cooler bodies. The warm water cools down and the cool water warms up. The process continues until a uniform temperature is reached at which time the temperature change ceases. It is quite a simple task using algebra and graphical analysis to find the quantitative relationship that relates the final temperature of the mixture to the initial temperatures of the cold and hot samples, respectively.

Level

Mainstream, first-year high school chemistry. Introductory level for all students.

Expected Student Background

See *Related Concepts and Related Skills*. Graphing skills and ability to transfer graphical information into an appropriate algebra expression are essential.

Time

Teacher preparation: 15 min

Student laboratory time: Approximately 50 min

Safety

Read the *Safety Considerations* in the *Student Version*. Normal safety precautions when using hot water should be observed. There are no chemical hazards.

Materials (For 24 students working in pairs)

24 Styrofoam® cups
24 Graduated cylinders, 100-mL
12 Thermometers calibrated in °C
12 Sections of rubber tubing to use with thermometer
24-36 Sheets of graph paper
12 Pencils
12 Rulers
Insulated coolers or coffee urn (2 or 3 for the class)

Advanced Preparation

Decide upon the type of hot water containers to be used and secure them before the laboratory time. Also, the teacher should remind the students to bring a calculator to the laboratory session.

Pre-Laboratory Discussion

By the nature of the exercise which uses Piagetian strategies, very little advanced discussion should be necessary. Just encourage the students to get to work on the activity and discover the existing relationships between the hot and cold water.

Teacher-Student Interaction

The teacher should circulate throughout the laboratory and assist the students in drawing and interpreting the graph of the hot, cold and mixture of water.

Anticipated Student Results

Some students will be unable to derive a relationship to find the final temperature of a mixture of hot and cold water from the data alone. In such instances, have these students make the graphs and find the areas before deriving a relationship or an algebraic formula. Some students may be able to derive a relationship from the data. In such cases have them make the graphs to confirm the formulas before you give them a set of data to check their algebraic formula.

Below is a sample of a set of data to check the algebraic formula.

Trial	Volume (Cold)	Temp (Cold)	Volume (Hot)	Temp (Hot)	Volume (Mix)	Temp (Mix)
1	30.0 mL	10.5 °C	49.0 mL	45.5 °C	79.0 mL	32.5 °C
2	40.0 mL	11.3 °C	40.0 mL	52.4 °C	80.0 mL	31.2 °C
3	50.0 mL	9.6 °C	50.0 mL	70.0 ° C	100.0 mL	39.6 °C

Actual temperatures of water kept in the insulated cooler or coffee urn can be quite different from student predictions based upon their laboratory experiment. This discrepancy is due to experimental design problems. Students often measure the hot water into a graduated cylinder directly and then pour it into the cup of cold water prior to taking the temperature. Consequently there may be considerable heat loss depending on delay of student work. This presents the instructor with an opportunity to discuss design of experiments. Several options are available depending on the instructor's inclination. One viable option is to have students brainstorm a new experimental procedure and repeat the laboratory on a subsequent day.

One procedure to produce better accuracy is to measure the volume of cold water with a graduated cylinder and add it to a Styrofoam® cup. Determine the temperature of the cold water and record it. Add hot water directly to the Styrofoam® cup and measure the temperature of the mixture after brief mixing. Be sure not to add too much hot water because the thermometer is covered beyond the 40 °C mark. Next, determine the final volume of the mixture using a graduated cylinder. Subtract the colder water volume from the mixture to find the hot water volume. Now substitute the variables into the algebraic equation and solve for temperature of the hotter water.

Answers to Data Analysis and Concept Development

1. The sum of the rectangular areas of the hotter and colder samples equals (approximately) the rectangular area of the hot and cold mixture.
2. $(V_c \times T_c) + (V_h \times T_h) = (V_c + V_h) \times T_f$
3. $(205\text{ mL})(30.0\ ^\circ\text{C}) + (105\text{ mL})(75.0\ ^\circ\text{C}) = (310\text{ mL})(\mathbf{X}) \quad \mathbf{X} = 45.2\ ^\circ\text{C}$
4. $\%\text{ error} = \dfrac{|\text{predicted temp - actual temp}|}{\text{actual temp}} \times 100\%$

   sample data: $\%\text{ error} = \dfrac{|68.0\ ^\circ\text{C} - 80.0\ ^\circ\text{C}|}{80.0\ ^\circ\text{C}} \times 100\% = 15\ \%$
5. Precision is usually good. The accuracy varies upon experimental design.
6. Answers will vary.
7. $(105\text{ mL})(15.0\ ^\circ\text{C}) + (70.0\text{ mL})(\mathbf{X}) = (175\text{ mL})(43.0\ ^\circ\text{C})\mathbf{X} = 85.0\ ^\circ\text{C}$

Post-Laboratory Activities

Compare student answers. Have the students brainstorm within their groups to determine how they can come closer to the temperature of the hot water that was determined by direct measurement. Have a contest to see which group can modify their experimental procedures to come the closest to the temperature in the coffee urn as taken by direct measurement.

Assessing Laboratory Learning

1. Challenge students to design an experiment to determine the temperature of the hot water by melting a known quantity of ice. Encourage them to do it as a home project and share their plan with their parents/guardians.
2. Assume an industry dumps hot (120 °F) water into a river. Your are part of a team to assess the impact on thermal pollution. Identify some of the parameters you are interested in measuring for the study?
3. Ducks are able to keep warm blood flowing from their feet back to their bodies. Vessels that carry blood from the body lie next to vessels that carry blood from the internal organs. Explain how this arrangement warms the blood to the body in the duck.

History Extension Supporting "Some Like It Hot"

Rumford and the Theory of Heat

Our knowledge of how solids and liquids behave is closely tied to our understanding of the kinetic theory of matter. Kinetic theory concepts depend heavily upon the nature of heat and heat flow. Until the early to middle nineteenth century, heat was believed to be an invisible fluid known as caloric. Caloric was thought to be squeezed out of little pockets in matter by the pressure of machines doing work on the surface of the matter. The influence of the caloric theory of heat can be seen yet today as heat is still said to "flow," and until very recently the accepted unit of heat was the calorie.

Much of the experimental evidence which led the the overthrow of the caloric theory was collected by Count Rumford. Count Rumford, an American expatriate, was in charge of boring cannons for the Bavarian army. Rumford found that so much heat was released during the boring of cannons that a constant flow of water over the barrel of the cannon was necessary to keep the cannon from becoming too hot and warping as the boring drill was doing its job.

To Rumford's surprise the cannon barrels heated up more when the drills were dull and very little actual material was being removed by the boring process. He determined this by measuring the amount of water necessary to keep the cannon cooled when boring with dull versus recently sharpened drills.

In an ingenious leap of reasoning, Rumford concluded that heat could not possibly be a fluid trapped in pockets in the cannon barrel. He reasoned that if caloric were a fluid trapped in the material, more of this invisible fluid would be released during rapid drilling. The sharpened drill would be swiftly cutting into new "pockets," freeing the caloric.

The fact that a dull drill produced more heat was inconsistent with the caloric theory. Rumford concluded that the dull drill could not be cutting into enough pockets of caloric to release the tremendous amounts of heat measured. He correctly reasoned that heat was a form of energy released by the friction as the surfaces were rubbed together. His theory was not fully accepted until James Prescott Joule measured the mechanical equivalent of heat in a series of experiments conducted nearly 50 years later.

References

Module drafted by David W. Brooks, Robert Curtright, and James McGahan, the Nebraska team.

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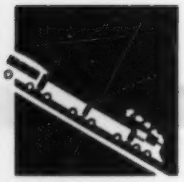
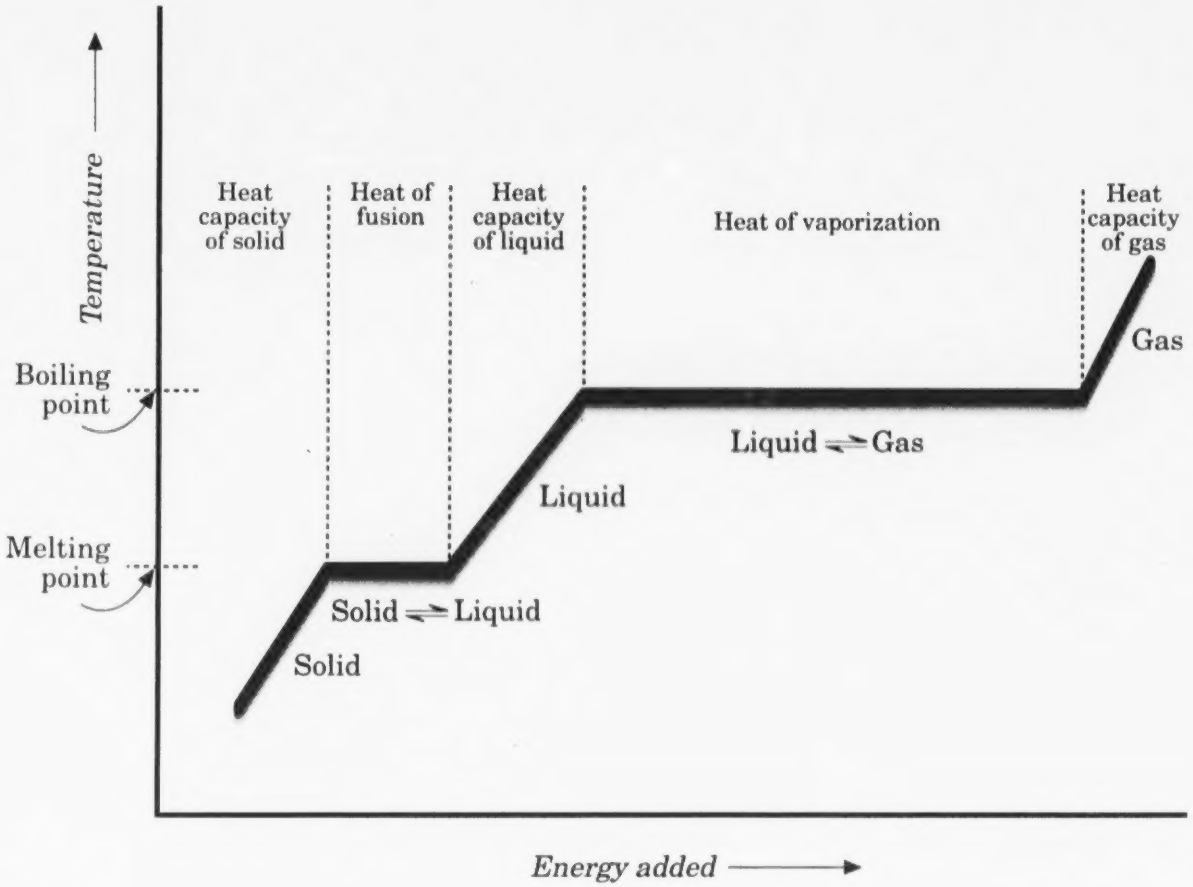
The polywater episode is an example of the self-correcting nature of science.



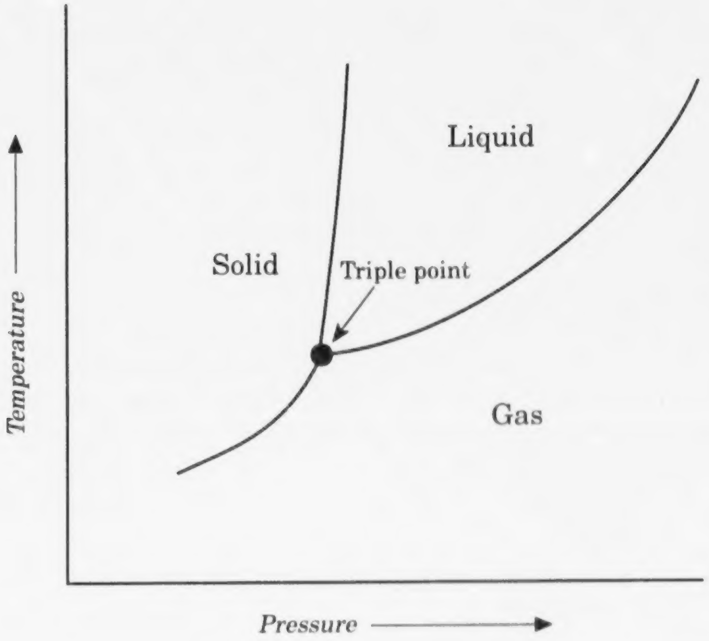
Appendix

- **Transparencies**
  1. Warming Curve
  2. Typical Phase diagram
  3. Word Search
- **Humor**

Warming Curve



Typical Phase Diagram

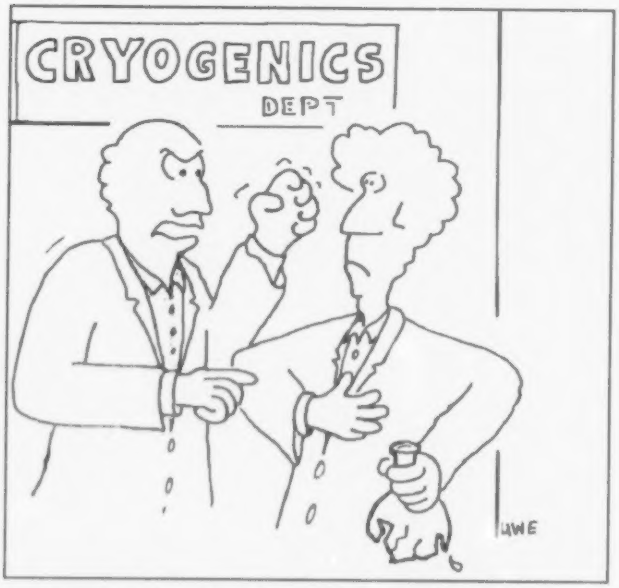

Liquid
Solid
Triple point
Gas
Temperature
Pressure

Word Search

L	A	T	E	N	T	H	E	A	T	B	Z	W	Z	W	N	B
W	N	N	U	R	Z	Y	P	X	E	O	M	T	E	S	M	H
I	E	I	H	Y	U	R	F	I	G	I	K	T	L	X	E	W
A	G	O	N	S	F	S	O	U	E	L	P	R	I	A	Q	D
E	O	P	V	K	G	T	S	U	G	I	C	J	T	A	D	E
M	R	G	I	C	Q	E	Y	E	N	N	F	C	A	M	Y	T
F	D	N	V	L	N	X	C	N	R	G	A	J	L	O	C	F
I	Y	I	T	N	O	I	S	N	A	P	X	E	O	E	B	L
F	H	T	S	O	Y	O	L	X	A	O	R	S	V	Q	S	S
F	T	L	B	R	O	S	Q	C	G	I	N	O	C	I	U	N
B	T	E	D	K	J	H	I	T	T	N	Y	P	P	D	J	O
E	W	M	N	O	I	T	C	A	R	T	N	O	C	A	U	W
S	O	N	M	P	Y	S	B	R	E	E	V	D	A	A	V	P

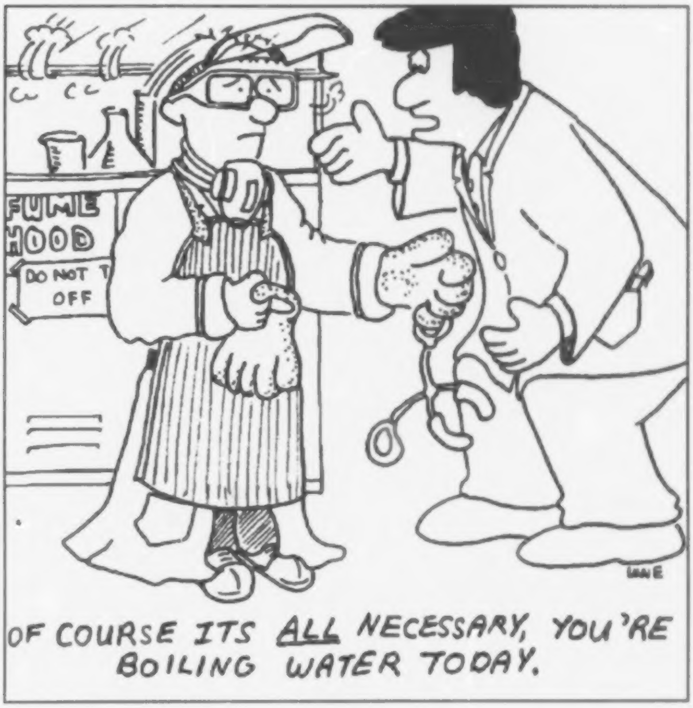
Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. This quantity is exerted by a gas that is in equilibrium with its liquid state at a given temperature. (2 words)
2. Temperature at which a liquid's vapor pressure equals the atmospheric pressure. (2 words)
3. Energy that is released but not apparent since it goes toward changing phase. (2 words)
4. The quantity of energy that a given object can absorb per degree Celsius change in temperature. (2 words)
5. What generally occurs when solid materials are heated due to increased kinetic energy of the particles.
6. Volume change that normally takes place when materials are cooled due to lower kinetic energies of the particles.
7. Type of bond that gives water its unusually high boiling point.
8. Adjective describing liquid that has a high vapor pressure and readily evaporates at room temperature.
9. The temperature at which a solid changes to its liquid form. (2 words)
10. Common name for $CO_2(s)$. (2 words)

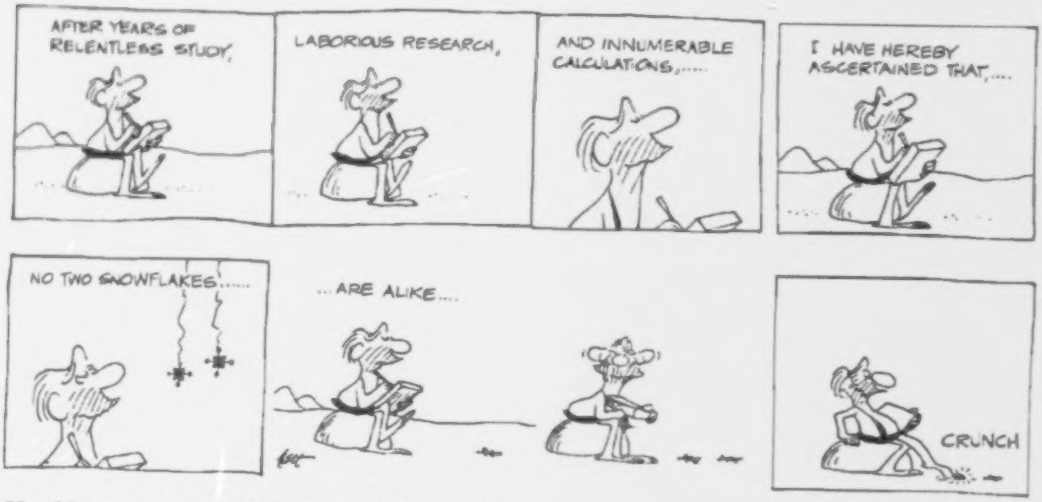

CRYOGENICS
DEPT
UWE

TEN YEARS OF RESEARCH AND YOU'RE THE ONLY ABSOLUTE ZERO WE'VE EVER HAD!

CHEM 13 NEWS, May 1980, p. 1


FUME HOOD
DO NOT T
OFF
LANE
OF COURSE ITS ALL NECESSARY, YOU'RE BOILING WATER TODAY.

CHEM 13 NEWS, May 1979, p. 17


AFTER YEARS OF RELENTLESS STUDY,
LABORIOUS RESEARCH,
AND INNUMERABLE CALCULATIONS,.....
I HAVE HEREBY ASCERTAINED THAT,....
NO TWO SNOWFLAKES......
...ARE ALIKE....
CRUNCH

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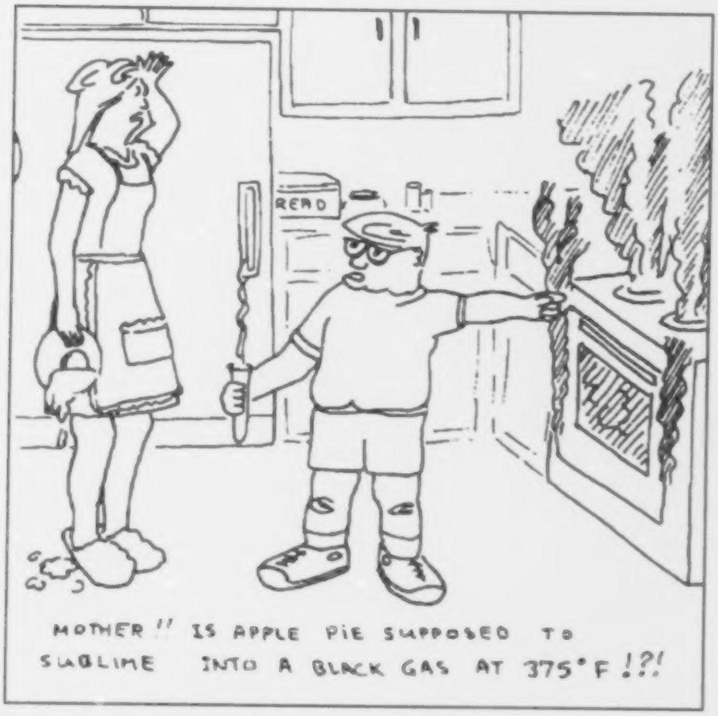
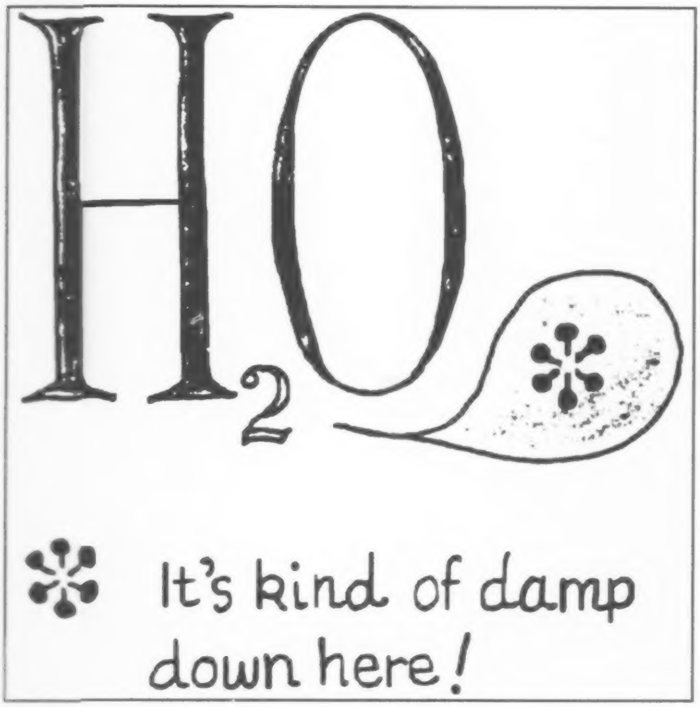

READ
MOTHER!! IS APPLE PIE SUPPOSED TO SUBLIME INTO A BLACK GAS AT 375°F !?!

CHEM 13 NEWS, April 1980, p. 5


H2O
It's kind of damp down here!

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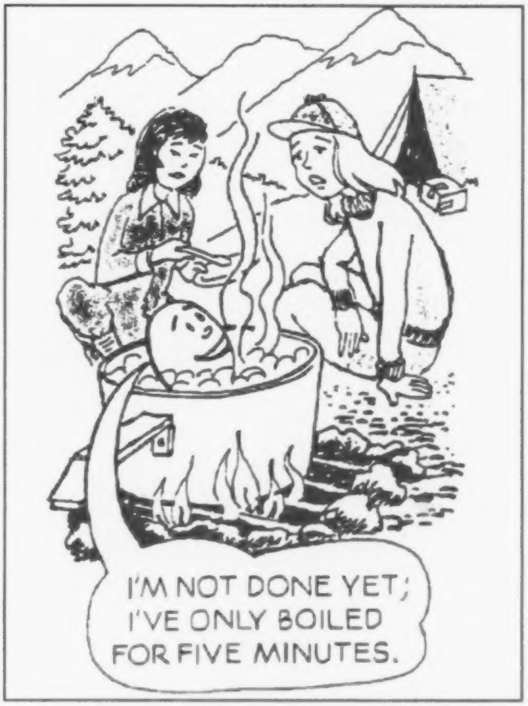
I'M NOT DONE YET; I'VE ONLY BOILED FOR FIVE MINUTES.

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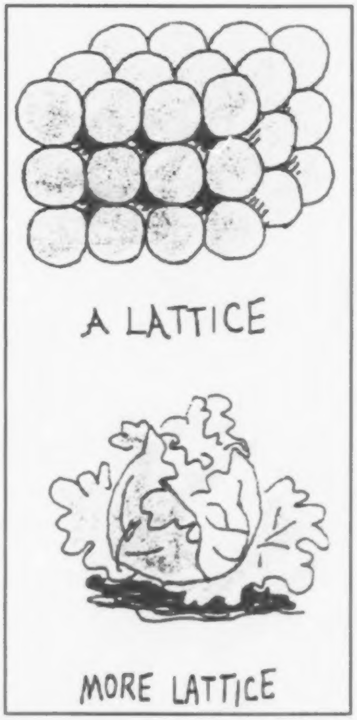
A LATTICE
MORE LATTICE

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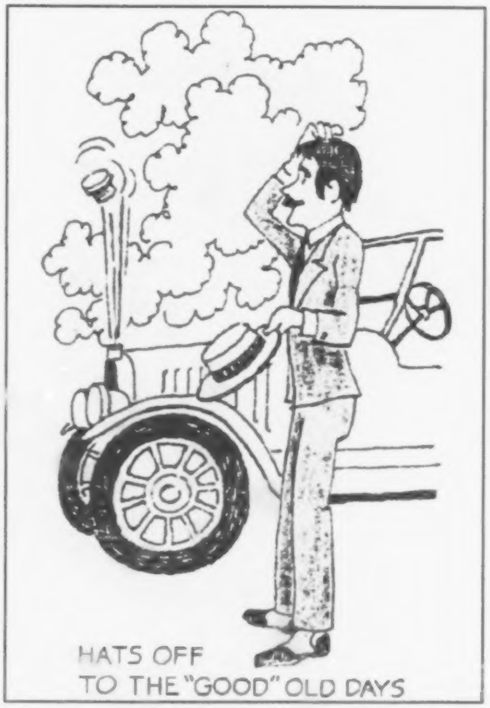
HATS OFF
TO THE "GOOD" OLD DAYS

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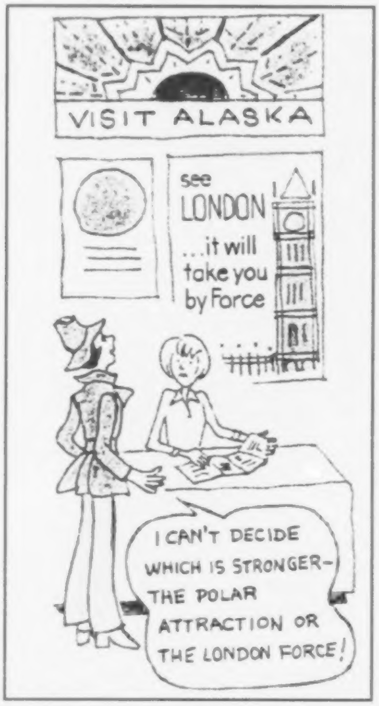
VISIT ALASKA
see LONDON ...it will take you by Force
I CAN'T DECIDE WHICH IS STRONGER– THE POLAR ATTRACTION OR THE LONDON FORCE!

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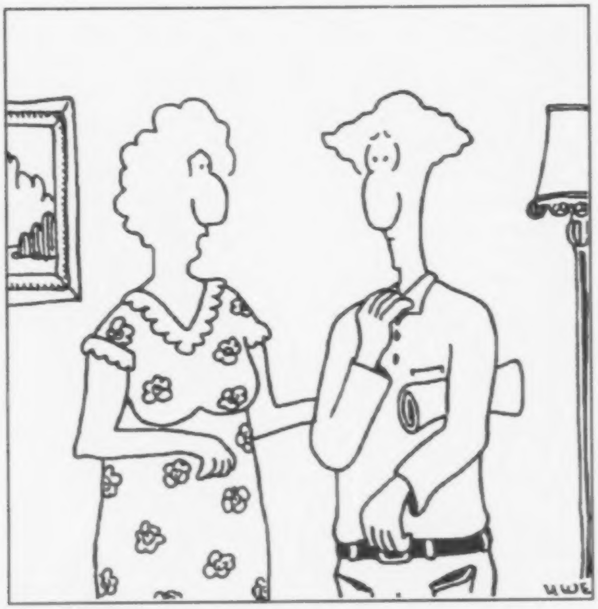
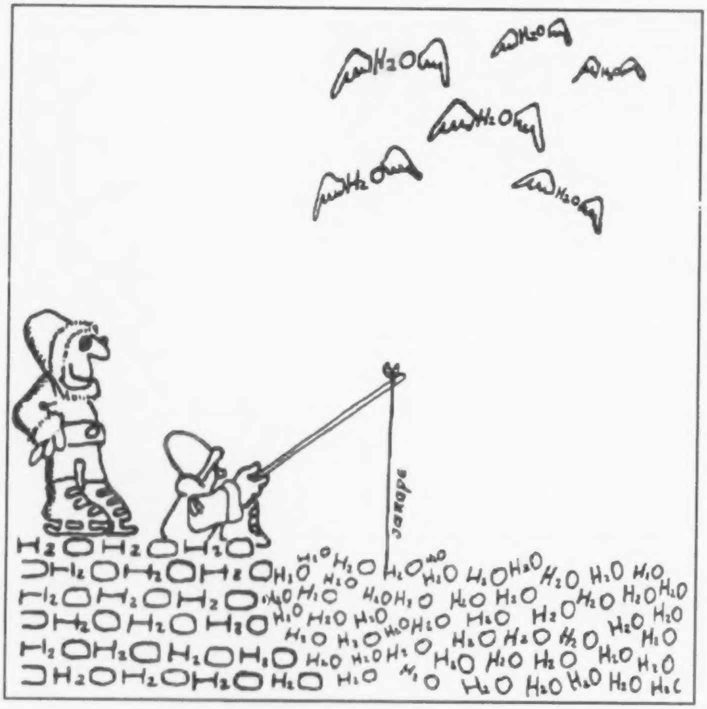


WHENEVER I MENTION WORK, LITTLE BILLY EVAPORATES OR MELTS AWAY. MAYBE IT'S A PHASE CHANGE HE'S GOING THROUGH.

CHEM 13 NEWS, February 1980, p. 2



CHEM 13 NEWS, September 1988, p. 27

A SourceBook Module

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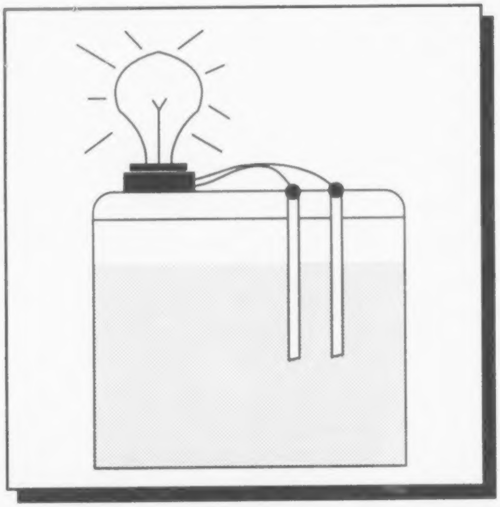
ELECTROCHEMISTRY



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Instructional Resources for Preservice and
Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

Like all of chemistry, electrochemistry is concerned with what electrons do. Electrons in atoms, molecules or ions are bound with a particular energy. When oxidation-reduction (redox) reactions occur, electrons are transferred from the substance being *oxidized* to the substance being *reduced*. In the process, energy is either released or absorbed, depending on the electron-binding energy difference between the reacting substances. A battery (a *voltaic cell*) is a device that allows the chemical energy released by a spontaneous oxidation-reduction reaction to do electrical work (*e.g.*, light a light bulb, power a radio.) An *electrolytic cell* can reverse this process by using external electrical energy to bring about a nonspontaneous redox reaction (*e.g.*, electroplating an automobile bumper.) In both cases, the electron transfer reactions occur at electrodes. The oxidation reaction occurs at the anode, and the reduction reaction occurs at the cathode.

Some substances give up or accept electrons easier than others and can be organized by their relative ability to undergo reduction reactions. A table of standard reduction potentials orders substances by their ability to accept electrons from a common donor (hydrogen gas). Among common substances, fluorine has the highest value (+2.87 V) for its reduction potential. It might be considered the *Tyrannosaurus Rex* of the Periodic Table because it can gobble electrons from anything. Lithium does not undergo reduction readily and has a very low reduction potential (–3.05 V). The difference between the reacting species in terms of reduction potential is an indication of the driving force for an electron-transfer reaction. In a voltaic cell, this difference defines the cell's potential and is the same as the electric potential for a standard single-cell battery. In an electrolytic cell, the cell potential is the electrical pressure (electric potential, expressed in volts) needed to drive the oxidation-reduction reaction.

Electrical charge is conducted through a solution by movement of cations and anions. In a voltaic cell, electrons are transferred from the cathode to the species being reduced. Cations must migrate to the cathode to offset the increase of negative charge near the electrode. Similarly, anions must migrate to the anode to offset the buildup of positive charge generated by the oxidation reaction at that electrode (see Transparencies 1 and 2 in the *Appendix*). The charge on the electrodes and the direction of movement of ions is just the opposite in electrolytic cells (seeTransparencies 3 and 4).

The quantity of work that can be produced by a battery or the quantity of work needed to run an electrolytic cell can be calculated from the electric potential, the current, and the time the cell operates. Thus, electrochemistry is the study of the interconversion of electrical and chemical energy.

PLACE IN THE CURRICULUM

The study of electrochemistry relies heavily on conceptual understanding of redox reactions. In essence, electrochemistry involves practical applications of redox chemistry. When students balance redox reactions by the *half-reaction method*, they are essentially separating the full reaction into two electrochemical half cells.

The structure and chemical properties of elements and molecular and ionic substances are governed by the chemical activity of their valence electrons. The energy that binds these electrons determines the value of a substance's standard reduction potential. Consequently, strong ties are apparent between electrochemistry and periodicity.

Whether an electrochemical cell operates in a voltaic or electrolytic mode depends on the relative free energy value for each half-cell. Observed phenomena in electrochemistry therefore directly relate to differences in thermodynamic functions of state (free energy, entropy, and enthalpy) and to temperature.

In electrochemistry, electrons are transferred in multiples of whole numbers. Therefore, there is a stoichiometric relationship between the quantity of charge transferred and the amounts of species oxidized and reduced.

CENTRAL CONCEPTS

1. Electrochemistry is a subtopic of oxidation/reduction.
2. Electrochemical cells contain several essential components:
    a. Cathode, where reduction occurs
    b. Anode, where oxidation occurs
    c. Electrolyte (can be liquid or gel; is sometimes a salt bridge), which serves to conduct charge *via* moving ions in the cell
    d. External circuit, where charge is conducted as a current of moving electrons
3. There are two kinds of electrochemical cells:
    a. **Voltaic** (also called Galvanic)
    Reactions spontaneous
    Chemical energy transformed to electrical energy
    b. **Electrolytic** (includes, but not limited to, electroplating)
    Reactions not spontaneous
    Electrical energy transformed to chemical energy
4. Electrochemical reactions occur at the surface of the electrodes.
5. Substances are categorized by half-reaction reduction potentials, which can be used to predict the spontaneity of oxidation-reduction reactions and the electric potential of batteries.
6. Technological applications of electrochemistry include batteries, fuel cells, electroplating, protection from corrosion, and chemical instrumentation.
7. Electrochemistry can be treated quantitatively at several levels of sophistication:
    a. Calculation of cell potentials for batteries (essential)
    b. Calculation of stoichiometric relationships (can be omitted if time is needed to treat other important chemistry topics)
    c. Electrical work (should be treated only if time permits and if students have completed physics as a prerequisite)
    d. Calculations using the Nernst equation or equations related to it or the Gibbs free energy equation (should not be treated in the general first-year high school course)

RELATED CONCEPTS

1. Electrical energy
2. Electric potential energy difference (volts)
3. Chemical energy
4. Spontaneous and nonspontaneous reactions

5. Conservation of charge
6. Anions, cations, and electrolytes
7. Electrostatic attraction/repulsion
8. Mole relationships in chemical reactions, balancing equations
9. Oxidation-reduction
10. Electronegativity, ionization energy, and electron affinity

RELATED SKILLS

1. **Math/computation skills**
   Basic arithmetic computation
   Graphing
   Determining all possible combinations
   Proportional reasoning
2. **Laboratory skills**
   Basic manipulation of glassware and chemical substances
   Completing circuits according to a circuit diagram
   Using a balance
   Reading meters and other instruments

PERFORMANCE OBJECTIVES

Upon completion of their study of electrochemistry, students will be able to:

1. distinguish between anions and cations.
2. define anode and cathode in terms of oxidation and reduction.
3. describe how a voltaic cell produces an internal ionic flow and an external electron flow.
4. write half-cell equations and total equations for voltaic cells.
5. in relation to a voltaic cell, define or explain: anode, cathode, electric potential (volts), salt bridge, internal circuit, external circuit.
6. describe how a battery produces electrical energy.
7. identify the substance being oxidized and the substance being reduced in an electrochemical cell.
8. describe the operation of an electrolytic cell.
9. define cathode, anode, and explain the charge on the cathode and anode in an electrolytic cell.
10. explain similarities and differences between voltaic and electrolytic cells.
11. explain the operation of an apparatus for electroplating with metals.
12. describe the zero-potential hydrogen half-cell.
13. given a table of standard reduction potentials: determine whether a redox reaction will occur; predict the electric potential of a voltaic cell made from two different half-cells; predict the products of an electrolysis reaction.
14. given the reduction potential of one half-cell and the electric potential of a voltaic cell, calculate the reduction potential of the other half-cell.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: A Study of Voltaic Cells

Purpose

To generate an activity series by constructing and comparing several electrochemical cells.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Any time you spill a solution on your clothes or body, wash it off immediately.
3. Silver nitrate spills on skin will cause dark stains and burns. Handle it with special care.
4. Dispose of all waste materials as your teacher directs.

Procedure

1. Collect the following materials:

   24-Well plate
   Salt-bridge
   Voltmeter (High-impedance)
   1-2 cm Strips of Zn, Mg, Ag, Cu, and Sn
   Dropper each of zinc sulfate, magnesium sulfate, silver nitrate, copper(II) sulfate, and tin(II) chloride

2. Examine the well-plate. Use any two adjacent wells to make a voltaic cell. Use the diagram below to design the most efficient arrangement of half-cells so you can measure the electric potential (voltage) of every pair of half-cells. To be most efficient, use one well that is not on the edge of the plate to make your silver half-cell. The silver half-cell involves expensive materials; you will assemble only one of these. If you are unsure about your arrangement, have your teacher check it before you start.

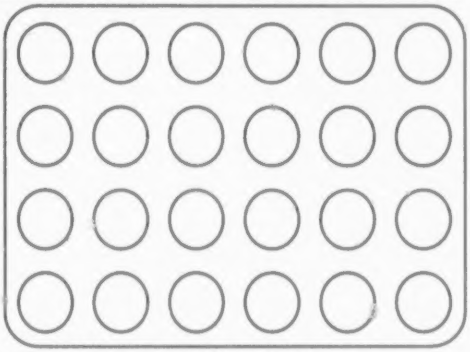


Figure 1. Well plate.

3. Fill each selected well with about 15 drops of appropriate metal ion solution.
4. Identify and label each of the five metals used in this activity according to directions provided by your teacher.
5. Clean each metal strip and wire with sandpaper. Place the cleaned metal samples on a piece of paper next to a chemical symbol that identifies each metal.
6. Make a table to record your data. For example, comparing Mg with Zn is the same as comparing Zn with Mg. You should find that 10 sets of measurements are needed to make all possible comparisons. However, if you are very clever and a little lucky, you'll be able to answer all the questions with a minimum of *four* comparisons.

7. Select two half-reactions to be tested. Connect them with a salt bridge. Be sure each end of the salt bridge is immersed in the half-cell solution.
8. Place appropriate metal strips or wires in the half-cell. For example, place Zn metal in the zinc sulfate solution, Ag in the silver nitrate solution, and so forth. *NOTE: Always use the correct metal electrode for each solution.*
9. Connect the metal strips or wires to the voltmeter. Attach the voltmeter so the cathode of the voltaic cell is attached to the positive lead of the voltmeter. You can tell by seeing which connection of the voltaic cell to the voltmeter gives a positive electric potential. Record the electric potential (in volts) for that pair, and record which metal is the cathode.
10. Continue testing pairs until you can answer all the questions your teacher has assigned.
11. Dispose of all solutions as your teacher instructs. The salt bridges and metal strips or wires can be reused; do not throw them away. Clean them as your teacher directs.
12. Thoroughly wash your hands before leaving the laboratory.

Data Analysis and Concept Development

1. Examine your data. List the half-reactions from highest reduction potential (most likely to be the cathode) to lowest reduction potential (least likely to be the cathode.)
2. If you have measured electric potential energy differences between pairs of half-cells and your teacher has provided the known reduction potential of oneof the half-reactions, calculate reduction potentials for the other four half-reactions.
3. Compare your results with a table of standard reduction potentials. Are there any differences between either the order you observed or the values you calculated? If there are differences, what might be possible reasons? (Hint: How were the standard reduction potentials measured, as compared with your procedure?) The reasons you offer to explain differences should be specific enough to be tested in the laboratory, if time allows.
4. Why must the salt bridge be in contact with both solutions for the voltaic cell to generate current?
5. Will zinc metal react when immersed in a solution of copper(II) sulfate? Will silver metal react when immersed in a solution of copper(II) sulfate?

Implications and Applications

1. Cathodic protection is the method most often employed to protect buried iron fuel tanks, iron pipelines, and iron ship hulls. An active metal that oxidizes more readily than iron is attached to the iron object being protected. Which metals used in this activity would be useful for the cathodic protection of iron? The reduction potential for the Fe/Fe^{2+} half-reaction is –0.41 V, while the reduction potential for the Sn/Sn^{2+} half-reaction is –0.14 V.
2. In 1973, the wreckage of the Civil War ironclad USS *Monitor* was discovered near Cape Hatteras, North Carolina. (The *Monitor* and the CSS *Virginia*, formerly the USS *Merrimack*, fought the first battle between iron-armored ships.) In 1987, investigations were begun to determine whether the ship could be salvaged. It was reported in *Time* (June 22, 1987) that scientists were considering attaching zinc anodes to the rapidly corroding *Monitor* metal hull. Describe how attaching zinc would protect the hull from further corrosion.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: A Study of Voltaic Cells

Introduction

Among several major types of chemical reactions, one of the most important is oxidation-reduction. You may have studied the electron transport system and photosynthesis in a biology course, for example. Both systems are just a series of oxidation-reduction reactions, each involving a species that loses electrons (is oxidized), while another species gains electrons (is reduced).

Even though oxidation cannot occur without reduction and vice versa, it is often useful to consider oxidation-reduction reactions in two parts called half-reactions. Added together, the two half-reactions make up the overall oxidation-reduction reaction.

$$M \rightarrow M^+ + e^- \quad \text{(Oxidation step)}$$

$$N^+ + e^- \rightarrow N \quad \text{(Reduction step)}$$

$$M + N^+ \rightarrow M^+ + N \quad \text{(Overall reaction)}$$

To decide whether a particular oxidation-reduction reaction will occur, it is helpful to think about the *reduction potential*. Reduction potential can be considered the driving force for a half-reaction to undergo reduction—that is, to go in the direction of the species gaining an electron. Consider a spontaneous oxidation-reduction reaction. The half-reaction with the larger reduction potential will occur as a reduction reaction. The half-reaction with the smaller reduction potential will run in the reverse direction—as an oxidation reaction in which electrons are lost rather than gained. For a more complete discussion of oxidation-reduction and half-reactions, refer to your textbook.

The standard way to compare half-reactions is in a voltaic cell. A voltaic cell is an electrochemical device that can produce electrical energy from spontaneous oxidation-reduction reactions. All electrochemical cells have two electrodes—a cathode and an anode. (An electrode supplies or accepts electrons from a chemical reaction.) Reduction reactions always occur at the cathode and oxidation reactions always occur at the anode. One easy way to avoid confusion is to remember that "**o**xidation" and "**a**node" both start with vowels, whereas "**r**eduction" and "**c**athode" both start with consonants. In voltaic cells, the cathode is charged positively and the anode is charged negatively. The identity of the cathode and anode is determined by the relative reduction potentials of the half-reactions that make up the voltaic cell. The electrode in the half-reaction with the more positive reduction potential is always the cathode in a voltaic cell. The electrode in the half-reaction with the less positive reduction potential is always the anode in a voltaic cell.

Unlike reactions that occur in test-tubes when substances are mixed together, voltaic cells are arranged so that half-reactions are physically separated, often into different containers called half-cells. The half-cells are connected by a conducting wire between the two electrodes. The conducting wire is often called the external circuit. An electrical current passes through the external circuit. Electrical quantities such as current (rate of flow of electrons), electric potential (electric potential energy difference between the two half cells), and resistance can be measured in this external circuit. The unit for measuring electric potential is the *volt*, thus electric potential is often referred to as *voltage*. The electricity flowing through the external circuit can be used to provide energy to systems such as flashlight lamps, audio recorders, and science experiments. A system in the external circuit is often called a load.

The half-cells are also connected by an internal circuit, often provided by a salt-bridge. The salt-bridge is composed of positive and negative ions that are free to move from one half-cell to the other but do not participate in the oxidation-reduction reaction. The salt-bridge is needed to keep charge from building up in the two half-cells. When electrons flow *from* a half-cell through the external circuit, negative ions travel to that half-cell to maintain electrical neutrality. Conversely, when electrons flow *into* the other half-cell, positive ions travel to that half-cell to ensure electrical neutrality there also.

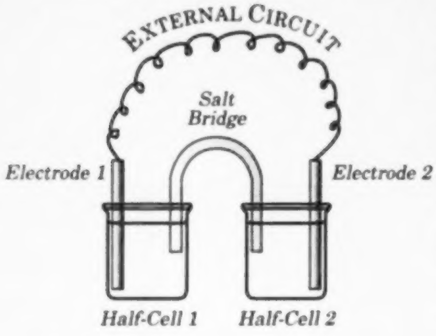


Figure 2. Voltaic cell.

Major Chemical Concepts

1. Electrochemistry represents a subset of oxidation-reduction.
2. Electrochemical cells include electrolytes, electrolyte bridge, electrodes labeled as anode and cathode, and an external circuit through which electrons flow.
3. Chemical energy is converted into electrical energy in a voltaic cell.
4. Half-cell potentials are relative, defined by arbitrary standards and an arbitrary zero value. (To successfully complete the activity and understand the concepts, it is not necessary to introduce the hydrogen half-cell as the arbitrary zero, but you may wish to do so.)
5. Charge is conserved in cells by the movement of ions.
6. Applications of the concepts developed in the activity include batteries and protection of artifacts from corrosion.
7. Cell potential or cell voltage—voltage associated with an electrochemical cell. It can be calculated from standard potentials and the Nernst equation.
8. The volt is the unit of electrical potential. It is defined as a Joule per Coulomb (1 V = 1 J/C).

Level

If you decide to have students *calculate* the half-reaction reduction potentials, this activity is most appropriate for general chemistry students. If students only measure the voltage, the activity may also be used with basic level students.

Expected Student Background

The following concepts are prerequisites for successfully understanding concepts developed in this activity: oxidation and reduction; the general nature of energy including potential energy, chemical energy, and electrical energy; the nature and behavior of charged particles, conductivity in ionic solutions; spontaneity of chemical reactions; arbitrary standards for measurement. Students should have developed fundamental laboratory skills. If you elect to have students calculate reduction potentials of the half-reactions tested, students should have arithmetic and/or calculator skills.

Time

If you prepare salt bridges, and dropper bottles of solutions in advance (see *Advance Preparation*), students should complete collecting data and clean up in 30-40 min. Very efficient students will be able to complete collecting data in considerably less time. Pre- and post-laboratory discussion times will vary, depending on how much concept development is tied to the laboratory experience.

Safety

See safety instructions in the student instructions. Similar laboratory activities in commercially available curricula often use lead or nickel half-reactions. Because lead compounds are toxic and some nickel compounds have been identified as carcinogens, we recommend that these *not* be used.

Materials (For 24 students working in pairs)

Twelve 24-well plates
12 Salt bridges (see *Advance Preparation*)
12 Strips of Zn, Mg, Ag, Cu, and Sn, 1-2 cm in length
0.1 M Zinc sulfate, $ZnSO_4$, 100 mL (2.9 g $ZnSO_4 \cdot 7H_2O$ per 100 mL solution)
0.1 M Magnesium sulfate, $MgSO_4 \cdot 7H_2O$, 100 mL (2.5 g $MgSO_4 \cdot 7\,H_2O$ per 100 mL solution)
0.1 M Silver nitrate, $AgNO_3$, 100 mL (1.7 g $AgNO_3$ per 100 mL solution)
0.1 M Copper sulfate, $CuSO_4 \cdot 5\,H_2O$, 100 mL (2.5 g $CuSO_4 \cdot 5\,H_2O$ per 100 mL solution)
0.1 M Tin(II) chloride, $SnCl_2 \cdot 2\,H_2O$, 100 mL (2.3 g $SnCl_2 \cdot 2\,H_2O$ per 100 mL solution)
0.5 M Copper(II) sulfate, $CuSO_4 \cdot 5\,H_2O$, 100 mL (12.5 g $CuSO_4 \cdot 5\,H_2O$ per 100 mL solution)
0.5 M Tin(II) chloride, $SnCl_2 \cdot 2\,H_2O$, 100 mL (11.3 g $SnCl_2 \cdot 2\,H_2O$ per 100 mL solution)
High-impedance voltmeter (see section on instrumentation below for suggestions)
Wires with alligator clips
Polarity tester (optional—see *Extensions and Projects*)

In addition to 0.1 M solutions for students, prepare 0.5 M and 1.0 M copper(II) and zinc sulfate solutions for post-laboratory discussion. Also have available some test-tubes and strips of Zn and Cu. Use the metals and solutions recommended. Not all combinations work well under the conditions of this laboratory procedure.

Advance Preparation

Check to ensure that all materials are available. Most materials are available in high school chemistry stockrooms. Silver wire can be ordered from a chemical supply company or may be available locally at a jeweler supply house.

Salt Bridges

Prepare in advance the salt bridges. Materials needed include potassium nitrate, burner, razor blade, and thin-stemmed polyethylene pipets.

Quickly wave the pipet through burner flame to warm it. (You will be surprised at how quickly the tube softens.) While it is still soft, bend the tube into a "U" shape and allow it to cool. Do not kink the tube while bending it. Use a razor blade to cut off the bulb and excess stem. Leave about 2 cm of tube on each "arm" of the bridge. Make all the "U" tubes you will need before proceeding to the next step.

Mix 2 g potassium nitrate in 10 mL of water. Add 0.1 g agar. Boil the solution for 3 to 5 min. While the agar is cooking, stretch the end of an uncut, unbent polyethylene pipet to reduce the diameter of its tube. Do this by warming in the burner flame and stretching. Cut the end of the pipet so it is even. Remove the agar from the heat. Use the modified pipet to fill each "U" tube with the warm agar solution. Be certain to eliminate all air pockets. After the salt bridge has cooled, trim the ends with the razor blade. Store the salt bridges in a jar containing 2 M potassium nitrate. The salt bridges will last indefinitely if they do not dry out. Make several extras; those used with silver and tin half-reactions will plug up after several uses due to precipitation of silver chloride. This can be avoided by substituting tin(II) nitrate or acetate for tin(II) chloride, but the former are not commonly available in high schools.

If you do not wish to make permanent salt bridges, filter paper strips or pieces of cotton twine soaked in potassium nitrate solution work fine as temporary and easily disposable salt bridges.

Pre-Laboratory Discussion

Pre-laboratory work will take two or three class periods if major concept development is tied to the activity. If your students have studied in advance all concepts to be developed by the activity, use the activity for concept review. The pre-laboratory would thus focus on demonstrating use of the apparatus. We recommend, however, that the activity be used for concept development.

Review prerequisite concepts briefly to ensure student understanding (see *Expected Student Background*). Ask certain students to explain each concept, followed by asking other students to elaborate, agree or disagree, correct, and provide examples. Your role is to serve as a "traffic director" for questions and responses, and to record on the board or overhead appropriate ideas to structure the review. You might administer an ungraded pretest quiz a day or so before the pre-laboratory discussion to help focus discussion. One useful way to do this is to provide a list of concept words in the domain being tested, such as oxidation-reduction or energy—and ask students to draw a concept map. (Concept maps are also useful after this discussion to summarize the review.)

Once prerequisite concepts have been reviewed, introduce laboratory equipment to be used. Demonstrate how voltmeter works by testing a small battery. Show students a 24-well plate with two solutions, metal wires or strips, and a salt bridge; connect the voltmeter.

If time permits, develop, *via* student discussion, the most efficient arrangement of wells for testing the various combinations and the minimum number of tests required for answering the questions. As students provide suggestions, ask them to explain how their arrangement will work and how data from their suggested tests can help answer the questions. Encourage students to propose tests that do not match up the silver and tin half-reactions. Although we recommend that *students* develop the well arrangement and test sequence, the arrangement in Figure 3 has been found to be very efficient. You can give it to students in advance to save time, but development of thinking skills will be compromised.

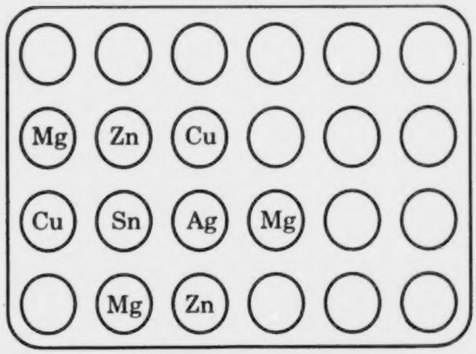


Figure 3. Well-plate arrangement.

Teacher-Student Interaction

As students conduct tests, have them identify the cathode and the anode, telling you what observations support their answers. Ask them whether oxidation or reduction is occurring in a particular half-cell, to define whether this is gain or loss of electrons, and to write the predicted half-reaction equation. You may wish to ask only one question of each group as you move from group to group. In that event, you might wish to prepare a checksheet matrix of question type by group so you can ask each group a conceptually different question each time you stop by.

Anticipated Student Results

Once laboratory errors such as placing the wrong metal into a well solution are taken into account, the half-reaction reduction potential order should be the same as that predicted by a standard table of reduction potentials. The actual values in volts for half reactions may differ somewhat from table values, depending on voltmeter impedance, meter reading errors, cleanliness of metal strip and wire surfaces, quality and size of salt bridge, and local concentrations of ions near the electrodes. Deviations from standard conditions under which tabled potentials are expressed should *not* much affect the results. (To understand why, review applications of the Nernst equation to cell potentials, as developed in college general chemistry textbooks.)

Answers to Questions

Data Analysis and Concept Development

1. See *Anticipated Student Results*.
2. See *Anticipated Student Results*.
3. Reasonable hypotheses include incorrect meter readings; using wrong metals in well plate solutions; and deviation from standard conditions of temperature, solution concentration, and/or pressure. Less likely to be proposed by students are surface coating of electrodes leading to changes in reduction potential, and lack of stirring in wells leading essentially to zero concentration of ions in well solutions near surfaces of electrodes. Accept any reasonable hypotheses as long as they are, in principle, testable.
4. For charge to be conserved in a half-cell, negative ions must replace lost electrons and positive ions must offset gained electrons.
5. Yes; No. (Be prepared to demonstrate this during post-laboratory discussion.)

Implications and Applications

1. Of metals tested in this activity, Zn and Mg can be used to protect Fe. For economic reasons, Mg is not used.
2. The Zn undergoes oxidation more easily than does Fe. (It has a larger negative—therefore lower—reduction potential.) Therefore, when Zn is attached to the Fe hull, Zn will corrode via oxidation rather than the Fe.

Post-Laboratory Discussion

Ask each group to report its results. Even if four of five groups obtain the same result, it is still possible that the *four* made a common mistake and the *one* completed the test correctly. Ask students to propose hypotheses to explain any discrepancies.

Ideally, student hypotheses can be tested directly at that time. A common hypothesis is that concentration makes a difference, so you should have available several solutions at higher concentrations to test that notion. By the close of this discussion, there should be general agreement on the observations—the cathode and anode in various combinations and the electric potential produced by those combinations—and reasons for any discrepancies. This discussion is an excellent opportunity to develop scientific thinking skills among your students.

Once observations are agreed on, process the information as a class. Repeat the questioning until an agreed order of reduction potential is established. If the activity's procedure is followed carefully, this order should agree with the sequence found in a table of standard reduction potentials. As part of this discussion ensure by questioning and explaining that students understand the concept of reduction potential.

After establishing the reduction potential order, give students the value from a standard table for one half-reaction used in the activity—any half-reaction potential will do. Once students know the half-cell potential for one half-cell, they can calculate the half-cell potential for another half-cell by finding the potential measured in the laboratory for that particular half-cell combination. Because a cell's electric potential is the sum of the potentials of its two half-cells, students can calculate a half-cell potential for the unknown half-cell. They can then proceed to calculate half-cell potentials for each half-cell studied in the activity. Our experience indicates that students learn to solve cell-potential problems much more readily by applying the relevant concepts to the laboratory *first*, followed by practicing textbook-type problems, rather than the reverse.

Extension

Have students use their text and school library resources to find out what materials are used in commercially available batteries. Then ask students to use the table of half-reaction reduction potentials to explain why many batteries provide 1.5 V of electric potential. Radio Shack sells a book explaining the makeup of many common types of batteries. In addition to exploring radio batteries, students can also calculate the electric potential of car batteries based on half-cell potentials and the number of cells in a car battery. Additional activities in commercially available laboratory books are excellent for reinforcing the concepts developed in this activity. "Experiment 46: Corrosion" (Wilbraham *et al.*, 1987) is highly recommended, as is the essentially identical activity from Carmichael and Haines (1987).

Assessing Laboratory Learning

Laboratory Practical

1. Provide students with two unknown metals and matching solutions. Ask them to determine which has the higher half-reaction reduction potential and/or the value of the cell electric potential. (The former task is more difficult conceptually.) The advantage of this assessment is its direct relation to laboratory procedures used by students. Disadvantages include set-up time, time needed for students to perform, and the large quantity of equipment required if individual students are to be tested simultaneously.
2. Provide students with a group of metals, test-tubes, and solutions. Without creating voltaic cells, have them determine the order, from lowest to highest, of half-reaction reduction potentials. Advantages include ease of set up, quick implementation by students, and the conceptual nature of the task. The primary disadvantage is that it does not duplicate the students' original laboratory procedure.

Oral Questions

If you wish, you can use questions recommended for teacher/student interactions during the activity as an evaluation tool. If you wish to assess individual student progress, make a student-by-question matrix to ensure that you ask each student the same questions—or at least the same number of questions. Then, as you circulate through the laboratory, ask each student a different question, checking their relative success. By circulating back through the laboratory several times, you can ask each student several questions before the end of the laboratory period, referring to your matrix to see which questions a particular student has been asked. If student performance on questions is to be used for grading, inform students in advance. We recommend that questions asked during the activity not be used for grading purposes.

Pencil and Paper

1. Provide students with written observations related to a similar laboratory activity. Ask students to interpret the information. **Advantage**: Duplicates the laboratory activity. **Disadvantage**: Requires considerable reading and therefore takes even good students a long time to complete.
2. Solve cell potential problems for half-cell potential, given the cell potential and the half-cell potential for one half-reaction. **Advantage**: Duplicates the laboratory activity. **Disadvantage**: Should probably not be used with basic level students.
3. Use questions recommended for teacher-student interactions during the laboratory activity later as written questions.

In the assessments described above, you may elect to allow students to refer to their notes, laboratory reports, and/or textbook.

Other Laboratory Activity Ideas

Give students some bubble solution made of about three parts good quality liquid detergent such as Joy and about one part glycerol, a 9-V battery, some connecting wires, and several pieces of aluminum foil. Have students electrolyze the soap solution with the Al electrodes, holding the electrodes close together so that bubbles form containing a mixture of hydrogen and oxygen gas. Holding a lighted match to the bubbles gives a satisfying but safe explosion. *[Edge, 1984]*

DEMONSTRATIONS

Demonstration 1: Voltaic Cells

Purpose

To demonstrate the ability of chemical reactions to produce electricity. The electric potential produced depends on the nature of the metal/metal ion half cell.

Materials

Petri dish
Filter paper
Dropper bottles of 0.1 M solutions of KNO_3, and Mg^{2+}, Zn^{2+}, Sn^{2+}, and Cu^{2+} as chloride or nitrate salt solutions.
Zn, Mg, Cu, and Sn strips, 1-2 cm
Wires with alligator clips
High-impedance voltmeter that will read 1-2 V

Safety

Use all normal safety precautions. Use accepted procedures for disposing of heavy-metal solutions.

Procedure

Cut the filter paper in the shape of a cross. Place it in the Petri dish. Add some KNO_3 solution to the center of the cross, allowing it to saturate the paper. Place one drop of each metal ion solution on each of the four corners. Attach two different metal electrodes to alligator clips wired to the voltmeter. Touch the proper electrode to its own metal ion solution. Substitute a different metal on the alligator clip until the electric potential of each of the six combinations is measured. This procedure avoids contamination and the need for a salt bridge, giving instant results.

Demonstration 2: Fruit and Vegetable Batteries

Purpose

To show that different combinations of electrodes produce different electric potentials and that some food materials can serve as electrolytes.

Materials

Metal strip of zinc, copper, aluminum, iron, tin, lead, 1 cm x 3 cm
Voltmeter
2 Connecting wires with clips on both ends
Firm fruit/vegetables (such as potato, lemon, grapefruit, orange)

Safety

Do not eat these materials. Wash hands thoroughly when the demonstration is completed.

Procedure

Cut the fruit sample in half and insert two unlike metal strips into it. Connect the strips to the voltmeter with connecting wires. Read the electric potential in volts. Experiment with placing the metal strips at various distances from each other. Try other fruit/vegetables with the same metals, and other combinations of metal strips. To reinforce concepts, have students predict in each case which of the two metals will be the cathode and which will be the anode.

Demonstration 3: Electrolysis of Water in Color

Purpose

To demonstrate electrode half-reactions.

Materials

Hoffman apparatus (or U-tube fitted with 2-hole stoppers and platinum; Nichrome electrodes; Bent wires in inverted test-tubes will also work.)
Bromthymol blue indicator (BTB)
6-V to 9-V DC source, such as a transistor battery
1 M HCl and 1 M NaOH, 5 mL each
Wood splints
2 Test-tubes

Safety

Follow normal laboratory safety precautions.

Procedure

Set up the Hoffman apparatus with its attached power source. Prepare the volume of water needed for electrolysis by adding approximately 3 g of solid sodium sulfate and a squirt of BTB. Turn on the power and observe the color changes during electrolysis. While electrolysis is occurring, show students the color of BTB in HCl (yellow) and NaOH (blue). After each arm of the apparatus has collected at least 20 mL of gas, note that the volume in one arm is about half that in the other arm. Note the color of BTB in each arm. Collect samples of gas generated at each electrode and test with glowing splints and burning splints to determine which arm contains hydrogen gas and which contains oxygen gas.

Reactions

Oxidation/anode	$6H_2O(l) \rightarrow O_2(g) + 4H_3O^+(aq) + 4e^-$
Reduction/cathode	$4H_2O(l) + 4e^- \rightarrow 2H_2(g) + 4OH^-(aq)$

NOTE: A U-tube set-up is preferred for seeing color differences, but withdrawing gas samples from a U-tube is difficult.

Demonstration 4: Electrolysis of Potassium Iodide

Purpose

To show electrode half-reactions involving color changes

Materials

Potassium iodide, KI, several crystals
Petri dish
Mechanical pencil leads
2 Connecting wires with clips on both ends
9-V Transistor battery
1% Phenolphthalein (in 95% ethanol)
Plastic sheet
Overhead projector

Safety

Use normal laboratory safety precautions. The iodine produces a brown stain on skin and clothing.

Procedure

Place a Petri dish on a plastic sheet on the overhead projector. Half-fill the dish with water and dissolve about a match-head quantity of KI in it. Make carbon electrodes by using the connecting wires to connect a pencil lead to each battery electrode. Arrange the two pencil leads so that only carbon is in contact with the KI solution in the Petri dish. Add one drop of phenolphthalein. Observe the two electrodes for evidence of a reaction.

Reactions

Oxidation/anode	$2I^-(aq) \rightarrow I_2(aq) + 2e^-$
Reduction/cathode	$2H_2O(l) + 2e^- \rightarrow 2OH^-(aq) + H_2(g)$

Phenolphthalein turns pink as OH^- collects near the cathode. The region near the anode turns brown as I_2 collects.

NOTE: You can also make an electrolysis apparatus for this reaction from a 9-V battery, two pencils, two connecting wires, a 2-hole stopper, and some tape. Sharpen two pencils and push them through the stopper. Near the erasers, carve the wood away until you can make contact with the "lead" inside. (You needn't carve wood from all around the pencil to do this.) Tape the battery to the pencils. Connect each battery terminal to the pencil "lead" near the eraser via connecting wires.

This system is much less fragile than the one involving two mechanical pencil leads, although it does take more time to make. You can attach the device to a ring stand with a universal clamp on the stopper to avoid holding it during the reaction. You can show the presence of I_2 *by squirting a little starch solution near the anode to show the blue-black color created by the starch test. (First demonstrate the starch test to your students with a known iodine sample.) If you use the pencil lead electrolysis apparatus with a solution of* $SnCl_2$*, you can show students Sn needles that "grow" on the cathode.*

Demonstration 5: Electroplating Copper

Purpose

To demonstrate industrial uses of electrolysis.

Materials

Copper strip, 1 cm x 5 cm
Stainless steel object (spoon, paper clip, knife, *etc.*)
18 M Sulfuric acid, H_2SO_4, 15 mL
Beaker, 250-mL
6 M Nitric acid, HNO_3, 50 mL
1.5-V Battery
2 Connecting wires with clips
Copper(II) sulfate, $CuSO_4$, saturated solution, 200 mL

Safety

Neutralize the acidic copper sulfate solution with an appropriate bicarbonate such as baking soda or a base like soda lime before disposing in accordance with local regulations. Dilute nitric acid can be flushed down the drain with ample water. Be careful to avoid contact with either acid. Copper salt solutions are toxic.

Procedure

Clean the copper strip by dipping it into a beaker of dilute nitric acid and washing it thoroughly. Prepare a plating solution by carefully adding 15 mL concentrated sulfuric acid *slowly with continuous stirring* to about 200 mL saturated copper(II) sulfate solution. Attach the copper strip to the positive battery terminal using a connecting wire. Attach the object to be plated to the negative battery terminal with the other connecting wire. Place the object to be plated and the copper strip in the solution. Rotate frequently for even coating. After 3-5 min, remove the plated object and observe.

Reactions

Oxidation/anode: $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
Reduction/cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

NOTE: An interesting reversal of the reaction can be demonstrated by switching the battery leads.

Demonstration 6: Making a Simple Battery: The Gerber Cell

Purpose

To demonstrate the ability of chemical reactions to produce electricity.

Materials

Large baby food jar
Dialysis tubing, 2-cm x 15-cm

1-Hole rubber stopper to fit jar
1 Strip each of copper and magnesium metal almost as long as the jar is deep and about half as wide as the dialysis tubing
Voltmeter
2 Connecting wires with clips on the end
0.5 M Copper(II) sulfate, $CuSO_4$, 100 mL (12.5 g $CuSO_4 \cdot 5H_2O$ per 100 mL solution)
0.5 M Sodium sulfate, Na_2SO_4, 100 mL (16.1 g $Na_2SO_4 \cdot 10H_2O$ per 100 mL solution)
6 M Hydrochloric acid, HCl, 50 mL
6 M Nitric acid, HNO_3, 50 mL

Safety

Use normal laboratory safety precautions and disposal procedures.

Procedure

Clean the Cu strip by dipping it briefly in dilute HNO_3 and then washing well with water. Clean the Mg strip by dipping it briefly in dilute HCl and then washing well with water. Fill the jar about 2/3 to 3/4 full of sodium sulfate solution. Wet the dialysis tubing and tie one end in a knot. Open the other end and fill to a depth slightly less than that of the jar with copper(II) sulfate solution. Place the copper strip in the dialysis tubing. Place the magnesium strip and the dialysis tubing in the jar. Then insert the stopper to hold them in place. Observe any reaction that may occur (see Figuure 4).

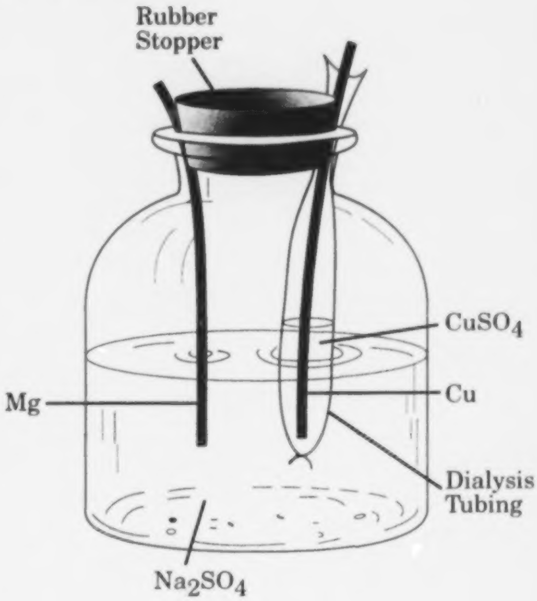


Figure 4. Simple battery apparatus.

Attach the ends of the Cu and Mg strips to the voltmeter with the connecting wires.

Reactions

Oxidation/anode: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$
Reduction/cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

NOTE: If you make six of these cells and hook them in series (anode to cathode), you can produce enough electric potential to operate a small radio that normally uses a 9-V battery.

Reference

Summerlin and Ealy. (1985). *Chemical demonstrations, Vol. 1.* Washington, DC: American Chemical Society.

GROUP AND DISCUSSION ACTIVITIES

Key Questions

1. What is the role of electrons in oxidation-reduction reactions? *[Electrons are the exchange particles of oxidation-reduction, moving from the oxidized to reduced species.]*
2. How can chemical reactions be used to produce electricity? *[If the half cells of a redox reaction are physically separated but connected by a salt bridge (to complete the internal circuit), electrons will flow through an external circuit of wire. The anion and electron movements are in opposite directions.]*

3. How can electricity be used to drive a chemical reaction? *[If a source of electric potential (such as a battery) is attached to electrodes in a solution or ionic melt, the applied potential forces electrons through the system. The electrode connected to the negative terminal holds an excess of electrons, attracts cations, and reduces them. This electrode is the cathode. Simultaneously, oxidation is driven at the anode as anions are attracted, lose their excess electrons to the electrode, and complete the electrical circuit. Electrolysis and electroplating processes rely on such reactions.]*

4. How do materials compare in their relative ability to accept or donate electrons? *[Materials tend to be classified as electron donors if they are metallic and receivers if they are nonmetallic. However, any particular reaction depends on the relative tendencies of the materials. Thus we rank materials according to their reduction potential. Even though hydrogen in contact with hydrogen ions is assigned a half-cell potential of zero volts, it (like other half cells) can either involve oxidation or reduction. The process observed depends on the nature of the other reactant. If it has a higher reduction potential, hydrogen molecules will be oxidized; if not, hydrogen ions will be reduced.]*

5. How are electrochemical processes used in business and industry? *[As sources of energy (electrochemical cells) and to drive chemical processes (electrolytic cells).]*

6. How are electrochemical devices used in everyday life? *[Applications of electrochemical cells and batteries (groups of cells in series) are myriad. Some examples include flashlights, portable appliances, pacemakers, automobile cranking systems, and submarine propulsion units.]*

Counterintuitive Examples/Discrepant Events

1. **Iron Can Be Protected from Corroding by Attaching It to Zinc**

   Iron is frequently used in construction and will corrode (oxidize) unless it is protected. One way to protect iron is to paint it. However, in many applications paint alone will not work because it is easily chipped or scratched. An effective way to protect iron is to attach a piece of zinc to it. This is frequently done to protect buried fuel tanks and hulls of ships. Zinc is easier to oxidize than iron and will lose its electrons (corrode) instead of iron. In such an application, zinc is a sacrificial cathode.

   Have students observe the corrosion of 6-penny iron nails. Start with two identical 6-penny nails. Clean the surfaces of both nails with a piece of sandpaper. Wrap one nail with 10 cm of zinc wire, making sure to establish good physical contact between the two metals. Immerse nails in a beaker containing a 10% sodium chloride solution. Observe the nails over several days. The bare nail will show signs of corrosion while the one wrapped in zinc will be protected from corrosion. (Students should recognize that galvanized construction materials are made of iron protected by zinc.) This activity is useful for initiating discussion about electrochemistry prior to Activity 1 in this module.

Analogies and Metaphors

1. Imagine an oxidation-reduction reaction as involving a hungry atom (draw a circle with chomping teeth in open mouth) eating an electron being fed to it by a nonhungry atom (closed mouth.) A voltaic cell just does the same thing, but the "hungry atom" has to be fed through a tube.

2. Use water analogies for electron transport.

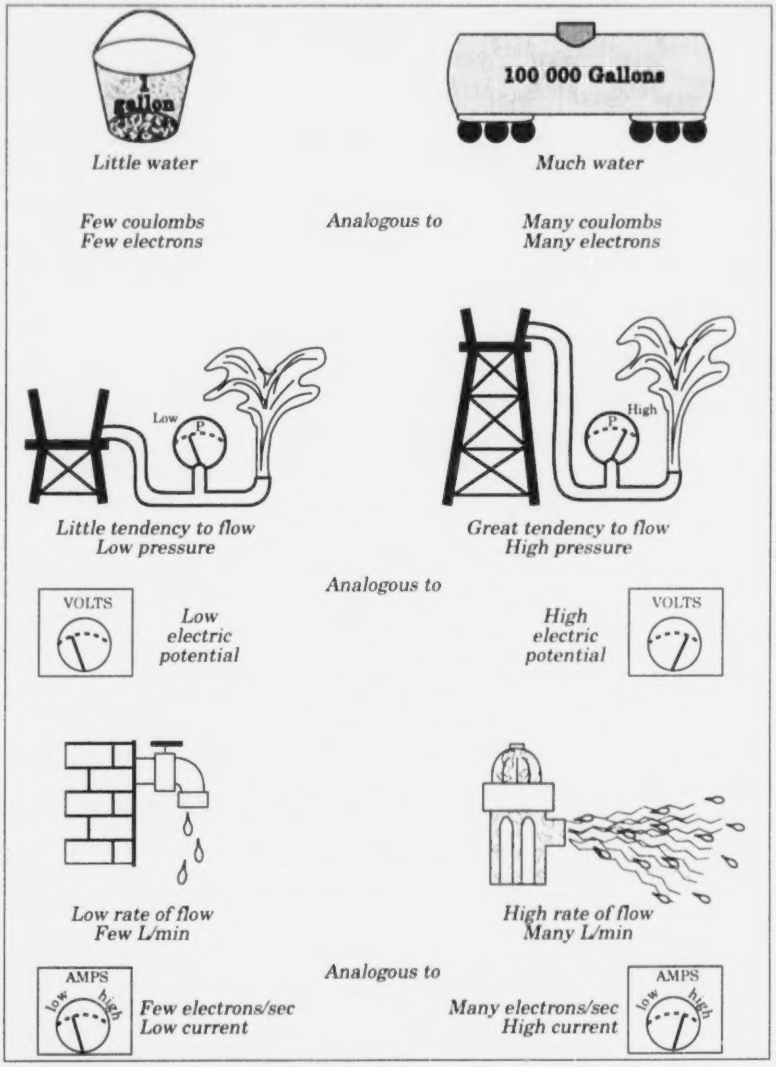

1 gallon
100 000 Gallons
Little water
Much water
Few coulombs
Few electrons
Analogous to
Many coulombs
Many electrons
Low
P
High
P
Little tendency to flow
Low pressure
Great tendency to flow
High pressure
Analogous to
VOLTS
Low electric potential
High electric potential
VOLTS
Low rate of flow
Few L/min
High rate of flow
Many L/min
Analogous to
AMPS
low
high
Few electrons/sec
Low current
Many electrons/sec
High current
AMPS
low
high

Figure 5. Electron transport/ water analogies.

3. Imagine that the quantity of work that can be performed by a galvanic cell is similar to the work obtainable from a water wheel powered by falling water.

The height of the waterfall is analogous to the cell potential. The higher the waterfall, the more potential energy it has. In the case of a voltaic cell, the higher the cell potential, the more the "driving force" for an electron. The quantity of water flowing over the waterfall is analogous to the current generated by the voltaic cell. The quantity of work that can be performed by a water wheel (equivalent to the voltaic cell) depends on both how much water flows each second (equivalent to current) and the height that water falls (equivalent to the electric potential.)

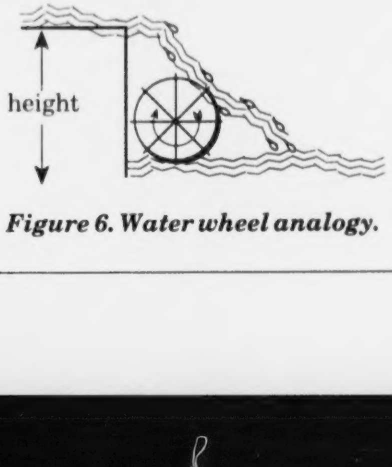


Figure 6. Water wheel analogy.

Other

Memory Aids: Mnemonics, *etc.*

1. Use a cat picture with + signs for eyes to remember Cation is positive. Or use cat paw with + sign in the pad print to remind students that cations are "pawsitive."
2. LEO the lion says GER. LEO = Loss of Electrons is Oxidation. GER = Gain of Electrons is Reduction.
3. Cathode and Reduction both start with consonants; Anode and Oxidation both start with vowels.
4. "An Ox and Red Cat" (ANode involves OXidation and REDuction involves CAThode)
5. OIL RIG = *Oxidation Is Loss (of electrons), Reduction Is Gain (of electrons)*

TIPS FOR THE TEACHER

Language of Electrochemistry

anode electrode at which an oxidation half-reaction occurs in an electrochemical cell

battery voltaic cell used to produce electricity through a chemical reaction

cathode electrode at which a reduction half-reaction occurs in an electrochemical cell

cell potential potential difference in volts of an electrochemical cell

cell voltage synonym for cell potential

corrosion oxidation of a metal by the action of air, water, and/or salt solutions

electric potential synonym for cell potential

electrochemical cell device in which electrons of a redox reaction pass through an electrical circuit

electrolytic cell electrochemical cell in which a nonspontaneous reaction is carried out by electrolysis

electromotive force (EMF) synonym for cell potential

galvanic cell electrochemical cell in which a spontaneous chemical reaction produces electricity

half-cell combination of reduced and oxidized forms of a given species upon which a redox equilibrium is established

nonspontaneous type of reaction when cell potential is negative

salt bridge device used to join two half-cells and containing a salt solution that permits the flow of ions between two half-cells

spontaneous type of reaction when cell potential is positive

voltaic cell synonym for galvanic cell

Common Student Misconceptions

1. **"Electrons can flow through solutions."**

   In "conduction of electricity" through solutions, electrons themselves do *not* pass through the solutions. Rather, charge balance is maintained in the solution by movement of cations and anions toward the electrodes where charge transfer takes place at the solution interface.

2. **"Water is a good conductor of electricity."**

   Water is a very poor "conductor" of electricity. (Recall that the ionization constant for water is very small.) The reason it is dangerous to insert a light bulb while standing in a puddle of water is that water is a great solvent for ionic compounds. Tap water and fresh water typically contain dissolved ions in sufficient concentrations to enable the solution to be conductive. However, ions in solution carry the charge and are thereby responsible for the current, not the water itself.

3. **"Only batteries are electrochemical cells."**

   Both voltaic and electrolytic cells are electrochemical cells. Voltaic cells produce electrical energy from differences in chemical potential energy. Electrolytic cells use electrical energy to produce products of higher chemical potential energy.

Problem Solving

1. Half-cell reduction potentials and voltaic cell potentials are intensive properties like density and temperature. Therefore, they are independent of the *quantity* of material present. That is the reason why multiplying half-reactions by a constant in balancing a redox equation does not affect the value of the potential for the reaction. Because the half-cell potential is an intensive property, it does not change when the amount of reacting material is changed. This is different from Hess's Law calculations and other calculations of thermodynamic quantities such as enthalpy, entropy, and free energy. As extensive properties, they do depend on the amount of substance present.
2. The numeric value for an oxidation potential is the negative of the reduction potential. When you reverse the equation for a reduction reaction, it becomes an oxidation equation and the direction (*i.e.*, the sign) of the potential changes.
3. When obtaining a reduction potential value in a standard table of half-reactions, the equation of interest must be identical to the one represented in the table. This includes phase (solid, liquid, gas, aqueous solution, *etc.*) as well as cases in which a precipitate might be involved such as the reduction of $AgCl(s)$ to $Ag(s)$ and $Cl^-(aq)$. Often, standard tables contain several equations that are very similar but have different half-cell potentials. Also, be aware that the reduction potentials in standard tables are calculated from thermodynamic quantities. Actual values measured in the laboratory may vary from these standard values. If you are interested in reasons, refer to the concept of *overvoltage* in a college chemistry text.

HISTORY: ON THE HUMAN SIDE

Michael Faraday (1791-1867). Michael Faraday was one of ten children of a London blacksmith and his wife. He had little formal education and was originally apprenticed to a bookbinder when he was 14 years old. Faraday taught himself to read in the bookbinder's shop and first learned about electricity, which was just being discovered during his teenage years, from an article in an encyclopedia brought to his employer for rebinding. Because electrical phenomena piqued his interest, he joined a group in London who attended lectures on science. He wrote complete notes on lectures he heard and had them bound in a book.

In 1812, he attended lectures given by Humphry Davy at the Royal Institution. He was so interested in Davy's talk that he applied to Davy, the institute director, for a job as his assistant. He finally obtained the position after following the advice of friends and sending his bound notes to Davy. This marked the start of Faraday's illustrious scientific career and his prolific collaboration with Davy. He went on to become the principal lecturer at the Institution and did much to popularize science among the British upper class; his Christmas lectures on the chemistry of candles remains a classic. In 1825, he replaced Davy as director of the Institution and his reputation soon began to rival that of Davy.

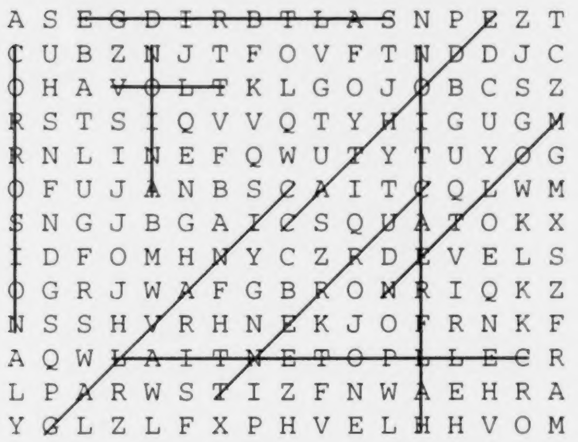
Faraday's work in analytical chemistry resulted in the discovery of benzene and many compounds of chlorine and carbon. He was the first to liquefy several gases, including CO_2, H_2S, HBr, and Cl_2. However, he is most famous for his work with electricity and electrochemistry. He discovered the principles that underlie modern electrical generators when he observed that a current is induced in a coil of wire rotating in a magnetic field. His experimental and theoretical work in electricity and magnetism formed the basis of later developments by James Clerk Maxwell. Albert Einstein rated Faraday along with Newton, Galileo, and Maxwell as one of the greatest physicists of all time. Faraday went on to prove that electricity generated in a magnetic field was identical to electricity produced by an electrochemical cell and performed many experiments measuring changes taking place in electrochemical cells. His work established two laws of electrochemistry still named after him in chemistry books: (1) The amount of material deposited in an electrochemical cell is proportional to the current passing through it, and (2) the amounts of substances deposited and/or dissolved in electrochemical cells are proportional to their molar masses.

A man of strong religious beliefs, Faraday tried to live a simple life, accepting rather reluctantly the many honors that came to him. His beliefs allowed him to solve without uncertainty a moral problem that still faces scientists. During the Crimean War between England and Russia in the 1850s, the British government asked him to head an investigation of the possibility of preparing large quantities of poison gas for battlefield use. Faraday refused to consider the idea, and nothing came of it at that time. Poisonous gas was eventually used in warfare during World War I in the early part of the 20th century.

Charles Hall (1863 - 1914) and Paul Heroult (1863 - 1914). The process used today for aluminum metal manufacture was invented independently at almost the same time by a young Frenchman, Paul Heroult, and a young American, Charles Hall. Hall invented the process while he was still an undergraduate at Oberlin College. He was inspired by a professor's remark that anyone inventing a cheap process for mass-producing aluminum would make a fortune. After graduation, Hall set up a laboratory in a woodshed—the 19th century equivalent to a garage—using homemade and borrowed equipment. After a year, he found that cryolite, Na_3AlF_6, would dissolve alumina (aluminum oxide) to give a conducting solution from which aluminum could be deposited by electrolysis. He used an iron frying pan as a container for the cryolite-alumina mixture, which he melted over a blacksmith's forge. The electric current came from electrochemical cells he made from jars used for canning fruit. (See the *Gerber Cell* demonstration described in this module.) As a result of the discovery made by Hall and Heroult, large-scale production of aluminum became economically feasible for the first time, and it became a common and familiar metal. Hall and Heroult share an interesting history. They were born in the same year, they discovered aluminum independently of each other in the same year, and they died in the same year.

HUMOR: ON THE FUN SIDE

1. Message on a T-shirt: I've got potential—Let's realize it!
2. "The patriotic ion went to the pole and volted." *(CHEM 13 NEWS, September 1980, p. 13)*
3. Word Search (see *Appendix* for master copy)



A S E G D I R B T L A S N P E Z T
C U B Z N J T F O V F T N D D J C
O H A V O L T K L G O J O B C S Z
R S T S I Q V V Q T Y H I G U G M
R N L I N E F Q W U T Y T U Y O G
O F U J A N B S C A I T C Q L W M
S N G J B G A I C S Q U A T O K X
I D F O M H N Y C Z R D E V E L S
O G R J W A F G B R O N R I Q K Z
N S S H V R H N E K J O F R N K F
A Q W L A I T N E T O P L L E C R
L P A R W S T I Z F N W A E H R A
Y O L Z L F X P H V E L H H V O M

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Driving force in a voltaic cell. (2 words)
2. Either of two parts of an oxidation-reduction reaction. (2 words)
3. Electrical quantity measured in amperes.
4. Unit of electrical potential.
5. Electrode at which reduction occurs.
6. Negative ion.
7. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction
8. Process by which metals are oxidized in the atmosphere
9. State in which NaCl must be in order to conduct electricity
10. Device containing an electrolyte that connects the two compartments of a voltaic cell. (2 words)

Answers: 1. CELL POTENTIAL 2. HALF REACTION 3. CURRENT 4. VOLT 5. CATHODE 6. ANION 7. GALVANIC 8. CORROSION 9. MOLTEN 10. SALT BRIDGE

4. Electrochemistry Crossword Puzzle (see *Appendix*)
5. See cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 15: The Busy Electron," can be used by chemistry teachers at all levels. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

2. *Electrical Interactions in Chemistry* is a film/video available from Ward's Natural Science Establishment, Inc., P.O. Box 92912, Rochester, NY 14692-9012; (800) 962-2660.

3. *Doing Chemistry*, videodisc set available from the American Chemical Society, 1155-16th St., N.W., Washington, DC 20036; (202) 872-4382.

   This three-disc set contains large numbers of laboratory and demonstration activities presented both as motion and still shots. It can be supported with Macintosh computer programs that provide access to HyperCard stacks for lesson planning and handout preparation, and that enable teachers to create their own videotapes. Over 700 pages of print materials are also available. In addition to the other fine activities, it includes two pertaining directly to electrochemistry: "DEMO E30: Daniell Cell;" and "DEMO E31: Galvanic Cell Based on Aluminum Oxidation."

4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).

   a. For the Apple II computer: AP 603 (includes Faraday 2, a simulation of Faraday's Laws in a laboratory experiment and Faraday Aid, a data analysis program used in conjunction with Faraday 2)

   b. For IBM PCs and PC-compatibles: PC 3101 (Faraday 2 and Faraday Aid)

5. Videodisc published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   "Copper-Zinc Electrochemical Cell," a chapter on *The World of Chemistry: Selected Demonstrations and Animations*: Disc I (double sided, 60 min.), Special Issue 3.

6. Software published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

   Frost Diagrams, by James P. Birk and Heidi Hocker. Vol. IV C, No. 1, for the Apple Macintosh.

EQUIPMENT AND INSTRUMENTATION

Equipment

1. **Power Supplies for Electrolysis.** Try using AC adapters used for tape recorders or transistor radios. They are rated for different electric potentials and currents, depending on their applications. 9-V, 500-milliampere adaptors commonly available at Radio Shack stores work well for most applications. Also, 9-V radio batteries will drive most aqueous electrolysis processes of interest in first-year chemistry courses.

2. **Supporting Electrolytes.** Sodium sulfate, potassium nitrate, and sodium nitrate are all good, unreactive electrolytes to use in electrochemical cells. Chloride salts will react with any silver cations formed, and will also undergo oxidation to chlorine gas at the anode of electrolytic cells.

3. **Salt Bridges.** Salt bridges can be made with a variety of methods and materials. See the laboratory activity section of this module for instructions on making agar salt bridges. A cotton string or piece of filter paper soaked with an electrolyte solution will work as long as it remains wet.

4. **Voltmeters.** If you attempt to measure the potential of a voltaic cell directly with a voltmeter, be sure to use a meter with high impedance. Many inexpensive meters, especially those sold in science education supply catalogs, are low impedance and draw too much current from the cell to accurately determine the cell potential. Radio Shack, Sears, and others market relatively inexpensive multimeters of high impedance that can also measure current and resistance.

5. **Inert Electrodes.** Platinum wire makes excellent inert electrodes. However, it is very expensive. A recipe for preparing platinum electrodes is provided under *Instrumentation*. Graphite pencil "leads" also work well, although they sometimes offer too much resistance if the pencil lead diameter is too small. Thick pencils used by young children and thick square pencils used by artists are easier to handle. Just carve the wood away from the graphite at both ends. Paint on the pencil wood can be removed by sanding. Old dry cell batteries usually contain carbon electrodes as well. Because they are much thicker than pencil leads, they do not break as easily and offer much less resistance. To obtain them, dissect old batteries. *(NOTE: Avoid alkaline, lithium, NiCad, mercury, and other such batteries.)* The carbon electrodes in the center can be cleaned with any strong oxidizer that has a lower reduction potential than manganese dioxide/permanganate half-reaction. The electrodes can also be sanded clean.

Instrumentation

Measuring the electric potential at which oxidation or reduction occurs under varying sets of conditions comprises a major subset of chemical instrumentation. You and your students can make simple electrochemical instruments for your laboratory. Directions for one such instrument are found in *Extensions and Projects*. Building any such instrument and making it work constitutes an excellent student project. The literature is full of references for building instruments. Specifically, look in *old* issues of the *Journal of Chemical Education*, "The Amateur Scientist" column in *Scientific American* when it was written by C. L. Strong, and *American Journal of Physics*. The electronics in those old instrument recipes are specified for vacuum tubes. These tubes can be duplicated by modern transistors and operational amplifiers. Find a ham radio operator, electrical engineer, electronic technician, or physics teacher to help you choose substitute parts. Instrument recipes as well as many other tips can be obtained at many science teacher meetings and conventions.

Links/Connections

WITHIN CHEMISTRY

1. **Stoichiometry and the Mole.** All electrochemical reactions are stoichiometric since an integral number of electrons must be transferred for each reacting atom, molecule, or ion. Many general chemistry texts include sections in which current in amperes and the time the current flows are used to calculate the total charge that has flowed through an electrochemical cell. This is, in turn, used to calculate moles of electrons transferred.

2. **Oxidation and Reduction: Commonly Used Batteries.** Dry cell batteries are a good example of the practical application of chemical principles. A standard design is to make the round "can" part of the battery of zinc metal and to serve as the anode. A carbon rod is inserted in the center to make contact with a mixture made of graphite carbon and manganese dioxide. This mixture serves as the cathode. An electrolyte/salt "bridge" mixture of moist zinc chloride, ammonium chloride, and nonconducting filler such as starch and flour surrounds the cathode mixture and connects it electrically to the zinc can. Not all features of the half-cell reactions are understood, although, in general, zinc loses electrons to become Zn^{2+} and manganese(IV) in manganese dioxide gains two electrons to become Mn^{2+} (see Figure 7).

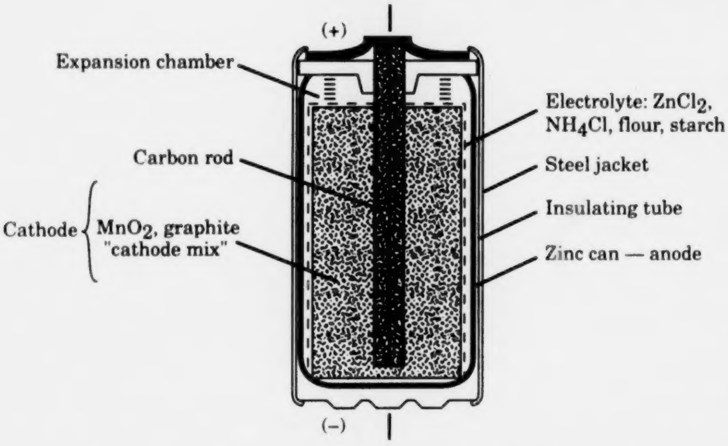

(+)
Expansion chamber
Carbon rod
Cathode
MnO_2, graphite "cathode mix"
Electrolyte: $ZnCl_2$, NH_4Cl, flour, starch
Steel jacket
Insulating tube
Zinc can — anode
(–)

Figure 7. Dry cell battery.

Alkaline batteries have numerous designs. A common "button"-type battery used for camera flashes and other small applications that demand high current uses mercury(II) oxide mixed with graphite as the cathode and powdered zinc amalgam (Zn and Hg) mixed with potassium hydroxide as the anode (see Figure 8). Because this type of battery contains mercury and mercury compounds, run-down batteries should be recycled to recover the mercury or treated to block entry of mercury compounds into the environment. Unfortunately, most people throw them unwittingly into the trash. A similar kind of battery uses silver oxide as the cathode and can be recycled for the silver.

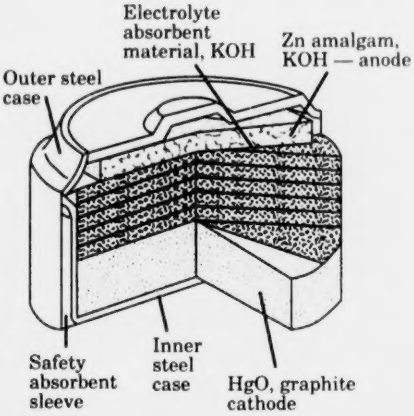

Electrolyte absorbent material, KOH
Zn amalgam, KOH — anode
Outer steel case
Safety absorbent sleeve
Inner steel case
HgO, graphite cathode

Figure 8. Alkaline "button"-type battery.

The general characteristics of several varieties of dry cells are given in Figure 9:

Substance oxidized	Electrolyte mixture	Substance reduced	Generic name
Zinc	Zinc chloride Ammonium chloride Filler, Carbon black	Manganese dioxide	Dry cell
Zinc	Potassium hydroxide	Silver oxide	Silver cell
Zinc	Potassium hydroxide	Manganese dioxide	Alkaline cell
Zinc	Potassium hydroxide	Mercury(II) oxide	Mercury cell
Cadmium	Potassium hydroxide	Nickel oxide	Ni/Cad (Rechargeable)
Lithium	Lithium iodide	Iodine complex	Lithium

Figure 9. Typical dry cells.

Details of lead-acid storage batteries for automobiles have been excluded from this module because explanations of this type of battery are included in most standard texts.

High-Tech and Advanced Batteries

For those who design batteries, a number of engineering performance parameters are important. Some of these are cost, performance at low or high temperatures, shelf life, lifetime in use, rechargeability, size, safety, energy or power density (quantity of current that can be delivered per unit time at a given electric potential), constant discharge electric potential (known as a "flat" discharge because when electric potential *vs.* time is graphed, the line is "flat" rather than negatively sloped), ease of construction, availability of materials (especially if the battery has military applications), and resistance of electrodes to corrosion and to forming an unreactive layer. Advances in batteries are typically made by improving materials from which electrodes and electrolytes (salt bridges) are constructed. One trick is to make a previously impractical material practical.

For example, sodium/sulfur batteries can potentially store about four times as much energy as the common lead-acid battery currently used in automobiles. A practical sodium/sulfur battery could make electric automobiles feasible. However, sodium/sulfur batteries operate at high temperatures and pressures. Therefore, the electrolyte material which separates them must be able to withstand those conditions while maintaining its ability to allow ions to migrate through it. Currently, researchers are attempting to develop a tough conductive ceramic by treating alumina with zirconium oxide doped with a small amount of rare earth oxide such as yttrium oxide. Other researchers are testing a sodium ion-containing glass made from oxides of silicon, sodium, aluminum, and zirconium.

Other systems under investigation as potential batteries for electric automobiles include zinc/bromine, lithium/iron disulfide, lithium/iron monosulfide, and aluminum/air batteries. Of those currently under development, only the aluminum/air battery has a theoretical energy density greater than gasoline, and is the only battery that could power an automobile for more than 300 miles without a recharge—the electrical equivalent of filling the tank. (The Ideal School Supply Company, Oak Lawn, IL 60453 sells an aluminum/air battery kit for educational purposes.) Scientists and engineers are constantly seeking new materials with interesting properties to accompany suitably high or low half-reaction reduction potentials. Other methods for improving batteries can include finding a coating for electrodes that allows them to function while protecting them from corrosion or the formation of an inert layer; redesigning the shapes of electrodes and their relationship to each other to lower internal resistance and to make recharging easier and more efficient; and spinning an electrode mechanically so that corrosion or surface layer formation is uniform.

Because batteries have so many applications and because the commercial potential of electric automobiles will be tremendous if an appropriate battery can be developed, large companies and entrepreneurial small companies invest heavily in electrochemistry research. For more information on advanced batteries see Brodd (1988) and Rawls (1985).

Fuel Cells

Fuel cells resemble voltaic cells in that electrochemical reactions occur in half cells connected by an external circuit. Batteries contain all the chemical reactants internally with only electricity flowing out (discharge) or in (during recharging). Fuel cells, however, constantly produce electricity but require reactants to be constantly fed in as "fuel." One simple fuel cell involves hydrogen fed into one electrode, oxygen fed into the other, producing electron flow through the external circuit and water.

Fuel cells are potentially much more efficient sources of electricity than are steam turbine-driven power plants. (See *Links to Physics*.) Several companies have built or are working on large fuel cells. Westinghouse Corporation has built a 1.5-MW fuel cell that uses phosphoric acid as the electrolyte and hydrogen and oxygen as reactants. It is called a PAFC Generator, where PAFC stands for Phosphoric Acid Fuel Cell.

Although fuel cells show promise as a system for generating electricity in the future, a number of technical problems must be solved. Most fuel cells use hydrogen as a fuel. A fuel cell that worked well with methyl alcohol would be more desirable because methyl alcohol is derived from plants more easily and cheaply than hydrogen gas can currently be produced. Many fuel cell combinations react too slowly to supply sufficient current for industrial uses. Catalysts that can speed up the electrode reactions are needed. Many noble metal (*e.g.*, Pt, Au, Ag, *etc.*) compounds work as fuel cell catalysts, but are expensive. Recent research has focused on using biological molecules such as porphyrins and phthalocyanines in conjunction with noble metal catalysts to increase the efficiency and decrease the amount of noble metal required. Fuel cells work best at high temperatures, but high temperatures are not possible for many applications; a polymer that conducts hydrogen ions (protons) at room temperature would be especially attractive as a solid electrolyte for many fuel cells (see *Potential Projects*).

One interesting application of the fuel cell idea is to turn a chemical plant into a fuel cell; for example, plants that make ammonia and methyl alcohol can be modified to become electricity-generating fuel cells. The opinion of many scientists and engineers is that fuel-cell technology will be a very profitable enterprise in the next few decades. For more information, see the 1988-89 series of articles by Lindstrom.

Electrochemistry and Dental Fillings

The material most commonly used for filling decaying teeth is called a dental amalgam. It is made by combining mercury with other metals such as silver or tin. If a person bites a piece of aluminum foil and the foil presses against a dental filling, a sharp, momentary pain often results. This result is due to a galvanic cell being created in the mouth. The aluminum serves as the anode and the filling is the cathode. Saliva acts as the electrolyte. A weak current flows between the electrodes and is detected by the sensitive nerve of the tooth as an unpleasant sensation.

Another type of reaction can occur if a filling makes contact with a gold inlay of a neighboring tooth. In this case, the dental filling made of mercury and tin acts as the anode and the gold inlay as the cathode. The dental filling starts to corrode. Prolonged corrosion will create the eventual need for a new filling. When this reaction occurs, tin(II) ions are released into the mouth, resulting in an unpleasant metallic taste. (See Transparency 5).

3. **Kinetics/Rates of Reactions**. One feature of electrochemical reactions is that the rate of reaction, unlike reactions in solution, is limited by how fast reacting molecules or ions can travel through solution to get to the electrode surface. When the external circuit is first connected, reactions occur very quickly involving molecules that happen to be close to the electrode at the time. Therefore, there is essentially a "burst" of current for a voltaic cell or a "burst" of electroplating or electrolysis for an electrolysis cell. Once atoms and molecules near the electrode are used up, reacting species must travel to the electrodes. If those reacting species are charged, they will be attracted to the charged electrodes. If the reacting species are not charged, travel to the electrodes will be controlled by the rate of diffusion unless the reaction mixture is vigorously stirred. Even with vigorous stirring, the layer, just a few molecules thick, near the electrodes might not be stirred.

   One way to stir the reaction at a molecular level if the species are charged is to place the electrochemical cell in a magnetic field, thereby forcing the ions to spiral toward the electrode, stirring the mixture on the way in. (This is analogous to the path taken by charged particles from the sun as they spiral into the earth at the poles under the influence of earth's magnetic field.) A suggestion for a student research project based on this principle appears in *Extensions and Projects*. Since the current that a voltaic cell or fuel cell can deliver depends on the rate at which reactants reach the electrodes, this is a very important effect. Once reacting species reach the electrode, their reaction rate depends on all the features of reactions in solutions. Catalysts play a role in batteries at the electrode surface, lowering the activation energy for and thereby speeding up the electron transfer reactions. The chemical trick to placing catalysts on the electrode surface is, of course, to maintain or even improve half-reaction reduction potential characteristics. For a good description of how electrochemists modify the surface of electrodes, read Faulkner (1984).

4. **Periodic Properties**. Although reduction potentials are more complex, they are similar in concept to the ideas of ionization potential and electronegativity, which are useful in discussions of periodicity. Where possible, periodic properties should be reinforced by comparing reduction potentials of elements within a single family.

5. **Producing Aluminum**. Aluminum is a very active metal and is therefore extracted from its compounds with difficulty. When aluminum metal first became available in the early 19th century, it was worth more than gold and other precious metals. All this changed with the advent of the Hall-Heroult electrolytic process, which is presently employed worldwide to produce commercial aluminum. The electrolysis is carried out in a carbon-lined vessel that serves as the cathode. The multiple anodes are made of graphite, and the electrolyte is purified bauxite (Al_2O_3) dissolved in cryolite, Na_3AlF_6. It requires 0.75 lb of graphite to produce 1.0 lb of aluminum, which underscores the importance of aluminum recycling. A large aluminum plant produces 4.0×10^8 kg (8.8×10^8 lb) of aluminum per year. It requires 200 kilowatt hours to produce 1000 kg (2.2×10^3 lb) of aluminum.
6. **Electroplating**. Electroplating is another important commercial process. Steel trash cans and steel siding are protected from rusting by electroplating a thin layer of zinc on the surface. A large silver anode, serving as a source of silver ions, is used to silver plate spoons, forks, and knives hung in a plating bath of silver cyanide. The plated objects serve as the cathodes in this electrolytic cell. A typical coating of silver has a thickness of about 0.1 mm.

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Fuel Cells and Carnot Cycle Heat Engines—Relationship to Physics**. One fascination of fuel cells (see discussion below) is that they are not heat engines like current fossil or nuclear fuel-based electric energy plants and automobile engines. Heat engines are all limited in their efficiency by the Carnot cycle; the difference between the operating temperature of the engine (typically a steam turbine for electricity generation or the inside of the cylinder for autos) and the environment (typically the outside world) divided by the temperature of the environment gives the maximum theoretical efficiency of a heat engine if everything works perfectly and there is no friction.

   For example, an automobile engine running at 627 °C (9000 K) on a hot day of 27 °C (300 K) would have an ideal maximum efficiency of 30%. Because real automobiles aren't frictionless or otherwise perfectly designed, the real fuel efficiency is less than 30%. This limit holds regardless of the kind of fuel used. By contrast, the efficiency of a fuel cell (and batteries, for that matter) are governed strictly by thermodynamic properties of the reactants—the electrochemical reaction's free energy change divided by its enthalpy change provides the theoretical efficiency of a fuel cell or battery.
2. **High-Tech Batteries/Fuel Cells and High-Temperature Super Conductors**. A challenge for many advanced battery systems researchers is finding materials that will work at high temperature and pressure. One approach is to make conductive ceramic electrolytes that will allow ions to move through them. Many rare earth and heavy metal oxides that are used to prepare high temperature superconductors (*e.g.*, yttrium oxide) are also used in preparing conductive ceramic electrolytes for advanced battery systems.
3. **Biology and Electrochemistry**. As mentioned in the introduction to the laboratory activities in this module, biochemical systems are electrochemical in nature. Two examples include photosynthesis and the respiratory chain of the mitochondrion. There are also electric eels, ion currents and action potentials in nerves, brain electricity as measured by the electroencephalogram, or EEG, and the ion currents in heart muscle measured by the electrocardiogram or EKG.

If your students are mathematically oriented, it might be interesting for them to calculate human power output based on the oxidation of glucose as a fuel as described by Chirpich (1975). An outstanding resource on bioelectricity and a number of other interesting topics is the book by Becker and Selden (1985).

4. **Modern Electrochemical Instrumentation**. Electrochemical instruments operate both in voltaic and electrolytic modes. Like all electrochemical cells, a minimum of two electrodes are employed. In some cases, a third half-cell is used as an internal reference cell; this half-cell determines the potential from which all voltages are measured. The most common reference cell is the calomel half-cell, sometimes referred to as the calomel electrode. The half-reaction for the calomel electrode is:

$$Hg_2Cl_2(s) + 2e^- \rightarrow 2Hg(l) + 2Cl^-(aq)$$

Electrochemical instrumentation is commonly used to determine the concentration of ionic species in aqueous solution. In the simplest case, an electrode is held at a constant potential *vs.* the reference. The potential is chosen so that an electrolysis reaction occurs that involves the species of unknown concentration. The number of electrons transferred at this electrode (the current) is related to the concentration of the species of interest. This technique is called polarography.

In addition, quantitative measurements can also be made in voltaic cells. The cell potential of a voltaic cell depends on the solution concentration of the species undergoing reaction at the electrode. (See a discussion of the Nernst equation in any standard college chemistry text for further information on this relationship.) Electrochemical measurements of pH and various other specific ion concentrations are examples of this voltaic technique.

TO THE CONTEMPORARY WORLD

Personal

Why an open flame is a hazard when charging an automobile battery. The typical auto battery is a lead acid battery in which sulfuric acid is the electrolyte. When the battery is charged, hydrogen gas can form at the cathode. This is not a big problem if you use a "trickle" charge (*i.e.*, low-current battery charger) in a well-ventilated space. Hydrogen gas is produced at a slow rate; ventilation allows most of it to escape, thereby preventing an explosive mixture. However a "fast" charge (*i.e.*, high-current battery charger) produces hydrogen at a fairly rapid rate. In a poorly ventilated area, the concentration of hydrogen gas can increase to the point where explosive mixtures are formed.

Community

1. **Field trips**. Electroplating companies, jewelry makers, electronic firms, chromeplating firms, battery companies
2. **Knowledgeable individuals**. Electroplating and electrolysis workers, individuals in battery industry, silversmiths
3. **Other community resources**. Local colleges/community colleges, any analytical chemistry laboratory using ion-specific electrodes

Extensions and Projects

Extensions

1. **Electrochemical gadgets**. A number of suppliers provide gadgets that can be used to illustrate electrochemical principles. For example, toy and novelty stores often have clocks that can be powered by a copper/zinc voltaic cell. A "two potato" clock that uses potatoes as electrolytes and salt bridges is often available. A company that sells interesting gadgets like this is The Home Scientist/Jerryco, 601 Linden Place, Evanston, IL 60202. Another is Grand Junction, Ltd., Science and Surplus Center, 100 South Lynn Shores, Virginia Beach, VA 23452. See also information in the *Demonstrations* section.
2. **Using school-built batteries**. A small transistor radio or "walkman" can be powered by several voltaic cells connected in series. Any number of "homemade" cells can be used as long as their potentials add up to the proper electric potential. Fairly large electrodes should be used to pass sufficient current.
3. **Electroplating unusual objects**. Electroplating with copper is a fairly easy process. If the object to be plated is not conductive, first coat it with a film of conductive material. There are commercial sprays available for this purpose. Hook up an electrolysis bath using the object to be plated as the cathode and a piece of copper sheet as the anode. The electrolyte solution should be about 1 M copper(II) sulfate. *NOTE: Copper does not plate well on a zinc object.*

Projects

Many chemistry teachers like to encourage students to complete research projects, either for their own edification or to enter into science fair and paper presentation research competitions. Below are a few suggestions for such projects.

1. **Gel electrophoresis**. *(This idea is by Sheila Kolb and Robert Roe, Jr., from the R. Arlotto et al. 1984 Dreyfus Institute Electrochemistry module.)* Electrophoresis is a method for separating mixtures that contain charges on molecules, or molecules with + and – charges like proteins, DNA, *etc.* The ions are attracted toward electrodes and forced to move through a medium such as gel. The combination of size, charge, and tendency to interact with the gel determine how fast various species travel and consequently how well and how quickly they can be separated from each other. Electrophoresis is not normally encountered in high school because of perceived expense and danger. Your local biology department may have an inexpensive unit. (Commercial firms use high electric potential.)

   However, your students can make electrophoresis separators with 9-V transistor batteries, a short piece of clear plastic tubing such as Tygon, some grocery store gelatin, and a small syringe and needle. The gelatin is used to fill a length of tubing. The pH of the gelatin can be controlled by gelling it with various buffer solutions made from acetate, phosphate, and ammonium salts instead of water. Dissolve about 10 g of gelatin in about 200 mL of water or buffer solution (a 250-mL beaker nearly full) in a hot water bath (a 400-mL beaker with enough water to form a "jacket" around the 250-mL beaker). Fill the tubing with gelatin and stopper at both ends until the gelatin cools and sets. Inject a small sample of the mixture to be separated through the tubing into the middle of the gelatin using the needle and syringe. A mixture of bromcresol green and methyl red is a good trial mixture for practice. Attach wires to the batteries and stick the ends of the wires in the ends of the gelatin.

Here are some questions that can be answered with this system.

1. What mixtures of ions can be separated? Can other mixtures such as amino acids or DNA fragments be separated like this?
2. What is the influence of pH on the separation?
3. What is the influence of concentration of ions in the water used for making the gelatin?
4. Does it matter how close together the electrodes are? Does the shape of the electrodes matter? For example, if you attached the wires to little flat electrodes that contacted the gelatin across the diameter of the tube, would the results be different?

2. **Making a Polymer that Conducts Hydronium Ions at Room Temperature.** Such a polymer was identified in *Links and Connections* as important for the economic development of fuel cells. Worthy (1985) reported news about such a polymer made from polyvinyl alcohol and phosphoric acid. One student might enjoy altering important variables such as reactant concentrations and preparation temperature to prepare such materials. The polymer is apparently made by mixing the two ingredients and evaporating the water. Details on testing the proton conductivity were not given in the article and might prove to be a challenge to a high school student. Before making a polymer, student(s) should look up the article and locate the research article by the original scientists by referring to Science Citation Index.

3. **Effects of Magnetic Fields on Electrochemical Reactions.** Ions moving in electrochemical cells are just moving charged particles subject to all the laws of physics including the one describing the force a magnetic field places on them. The rates and perhaps other features of electrochemical reactions involving ions should therefore be affected by placing them in magnetic fields. Test-tube reactors placed between strong magnets such as those available from many science supply companies are good systems for testing these effects. Students can also build Helmholtz coils and, with a variable electric potential source, can even subject the electrochemical cells to magnetic fields of different strengths. If you are unfamiliar with Helmholtz coils, consult a high school or college physics teacher. One easy system to measure is electrolysis of water. How does the presence of the magnetic field affect the rate at which gas collects in a tube?

Polarity Tester

A polarity tester is constructed of materials listed in Figure 10. It is used to determine which electrode in an electrochemical cell is the cathode and which is the anode.

Quantity	Description	Radio Shack #
1	Operational amplifier (Op Amp) Type 741	276-007
1	Light-emitting diode	276-226
1	Signal diode IN 914/4148	276-1620
2	9 V Battery snap connector	270-325
2	Mini gator clips	270-374
1	Op amp socket	276-1995
2	4.7 k-ohm 1/4 W resistor	271-1330
1	150 k-ohm 1/4 W resistor	271-047
1	560 ohm 1/4 W resistor	271-020
1	Circuit board, 24 gauge black and red wire	276-159A

Figure 10. Materials for a polarity tester.

Use Figure 11 as a guide in construction. Care is needed to not overheat any of the elements and to not cause a short by using too much solder in connections. Students interested in electronics may be induced to construct the polarity testers for you. If your school has an electronics course or a vocational education department, they might also help you build one. Once you have built a polarity tester or two, they take about 15-20 min to build. They are quite durable and should last many years with proper care.

Inexpensive multimeters of high impedance can also be purchased at many electronics stores (such as Radio Shack). These are good alternatives if you do not have the time to make your own polarity testers.

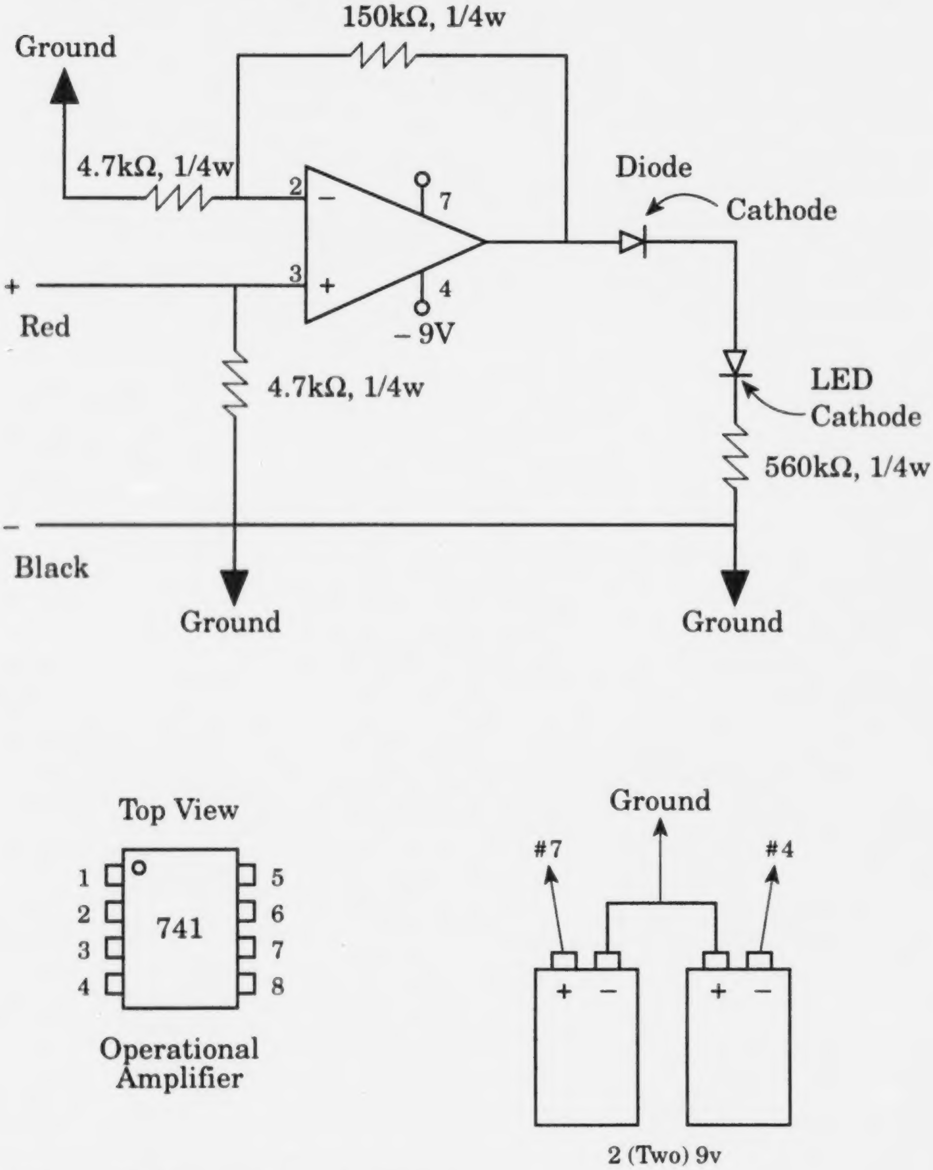

150kΩ, 1/4w
Ground
4.7kΩ, 1/4w
2
–
7
Diode
Cathode
+
3
+
Red
4
– 9V
4.7kΩ, 1/4w
LED
Cathode
560kΩ, 1/4w
–
Black
Ground
Ground
Top View
1
2
3
4
741
5
6
7
8
Operational Amplifier
Ground
#7
#4
+ –
+ –
2 (Two) 9v

Figure 11. Diagram for polarity tester.

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Module developed by William Becker, Beatrice E. Epperson, and William Lamb, the Oregon team.

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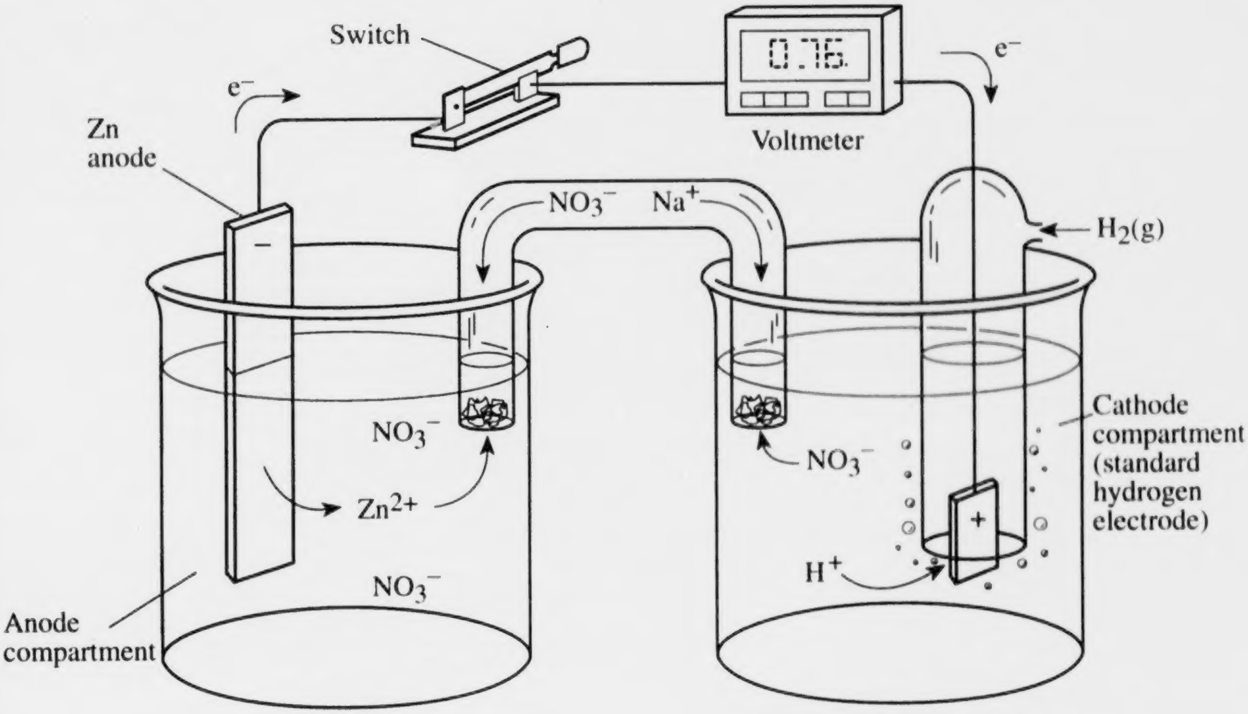
Appendix

- **Transparency Masters**
  1. Voltaic Cell: Zinc/Hydrogen Electrodes
  2. Voltaic Cell: Zinc/Copper Electrodes
  3. Electrolysis of Molten Sodium Chloride
  4. Electrolysis of Water
  5. Electrochemistry and Dental Fillings
  6. Word Search
  7. Electrochemistry Crossword Puzzle
- **Humor**

Voltaic Cell: Zinc/Hydrogen Electrodes

$$E^\circ_{\text{cell}} = E^\circ_{\text{ox}} + E^\circ_{\text{red}}$$

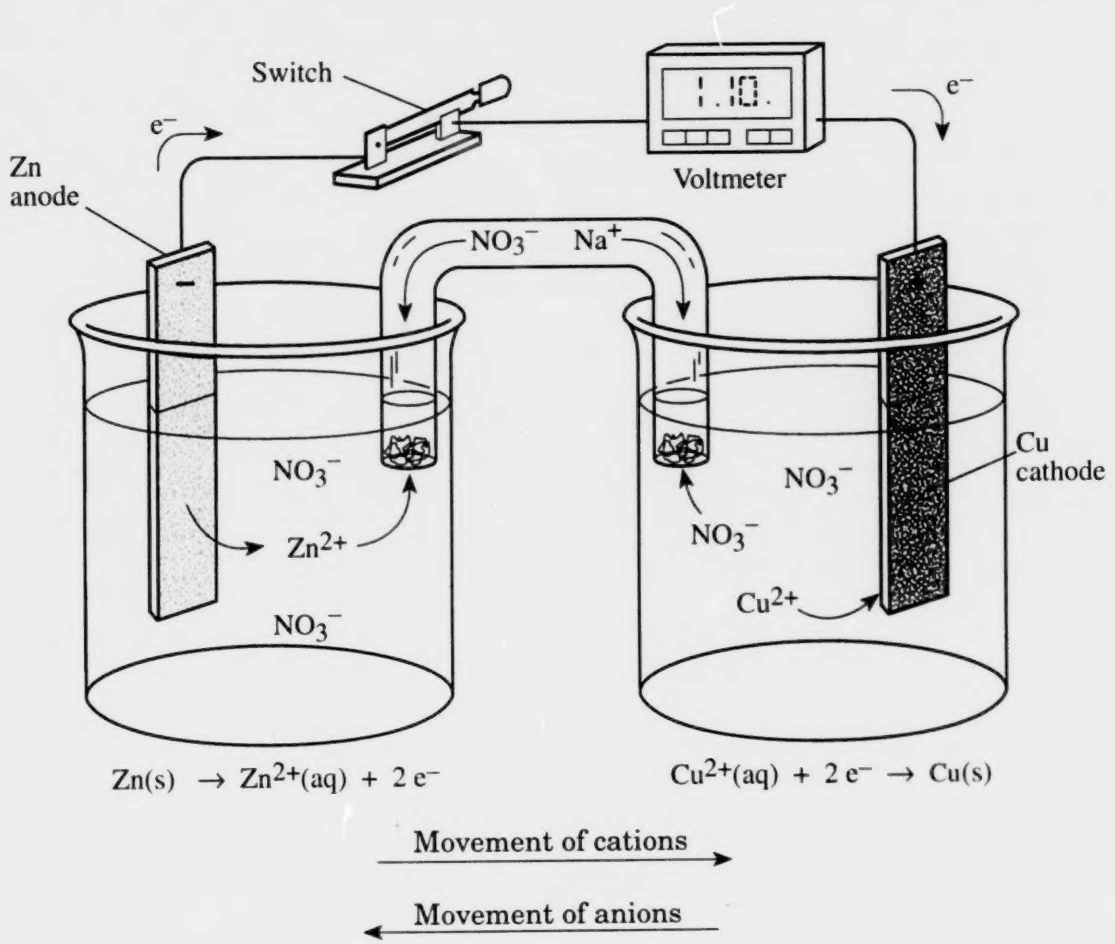
$$0.76\text{V} = E^\circ_{\text{ox}} + 0$$


Switch
0.76
e⁻
e⁻
Zn anode
Voltmeter
NO₃⁻
Na⁺
H₂(g)
−
NO₃⁻
NO₃⁻
Zn²⁺
Cathode compartment (standard hydrogen electrode)
+
H⁺
NO₃⁻
Anode compartment

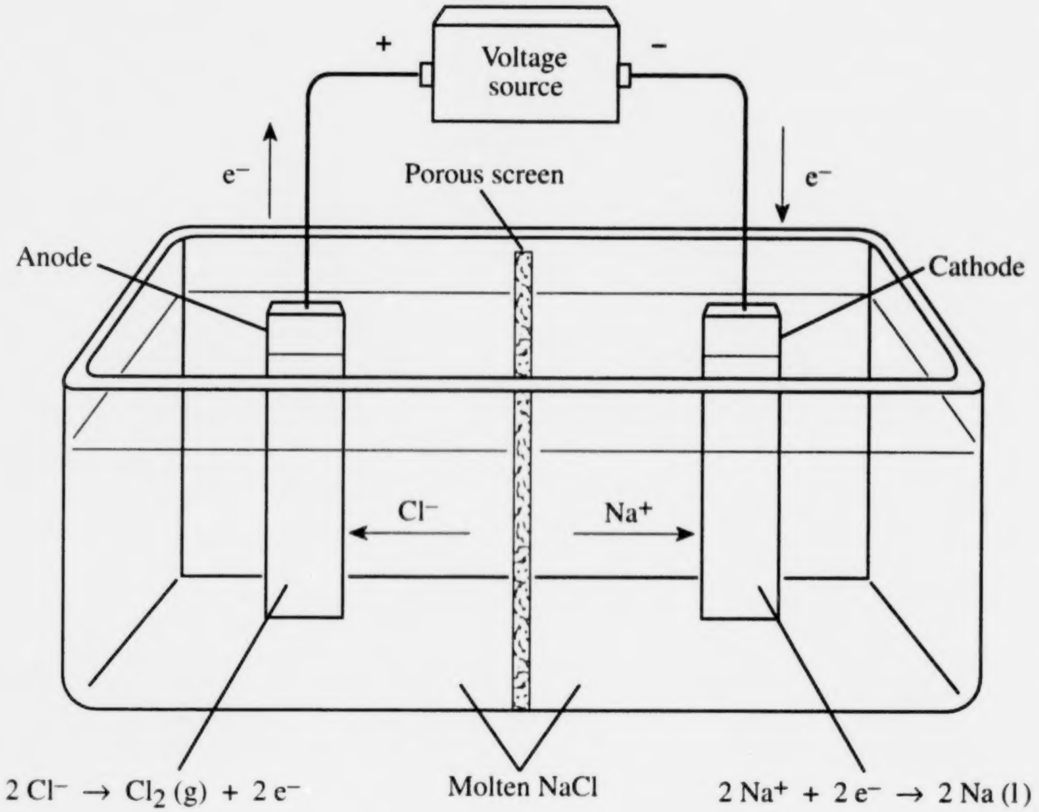
$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$

$2H^+(aq) + 2\,e^- \rightarrow H_2(g)$

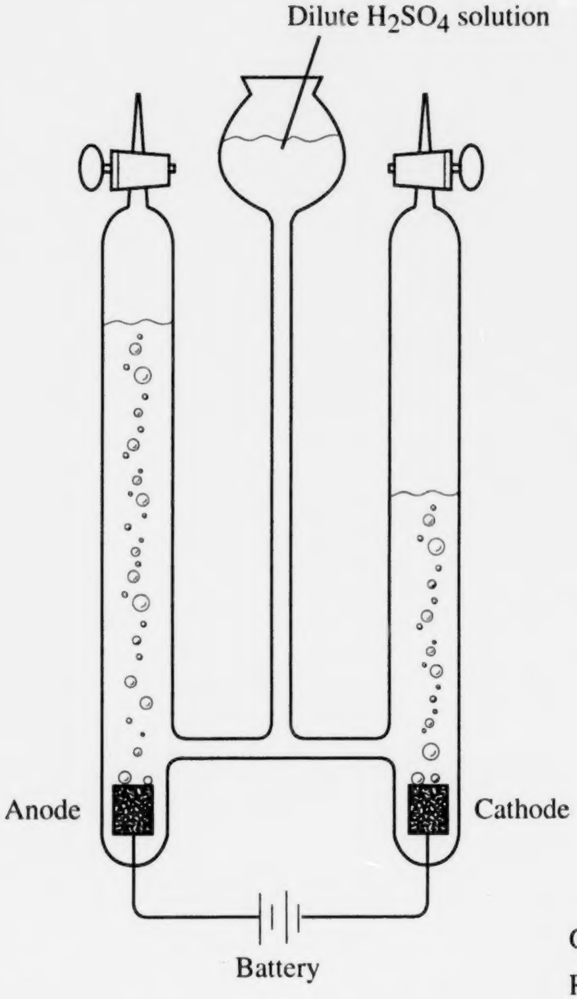
Voltaic Cell: Zinc/Copper Electrodes

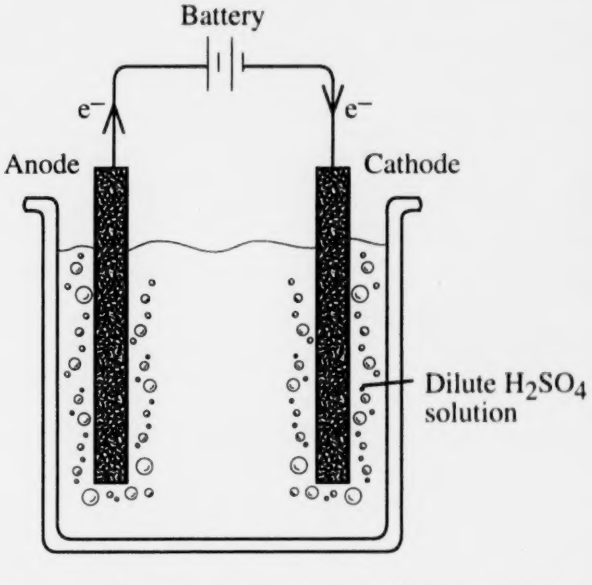

Switch
1.10
Voltmeter
e^-
e^-
Zn
anode
−
NO_3^-
Na^+
NO_3^-
Zn^{2+}
NO_3^-
Cu
cathode
NO_3^-
NO_3^-
Cu^{2+}
$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$
$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$
Movement of cations
Movement of anions

Electrolysis of Molten Sodium Chloride


+
Voltage source
-
e⁻
Porous screen
e⁻
Anode
Cathode
Cl⁻
Na⁺
$2\,Cl^- \rightarrow Cl_2\,(g) + 2\,e^-$
Molten NaCl
$2\,Na^+ + 2\,e^- \rightarrow 2\,Na\,(l)$

Electrolysis of Water

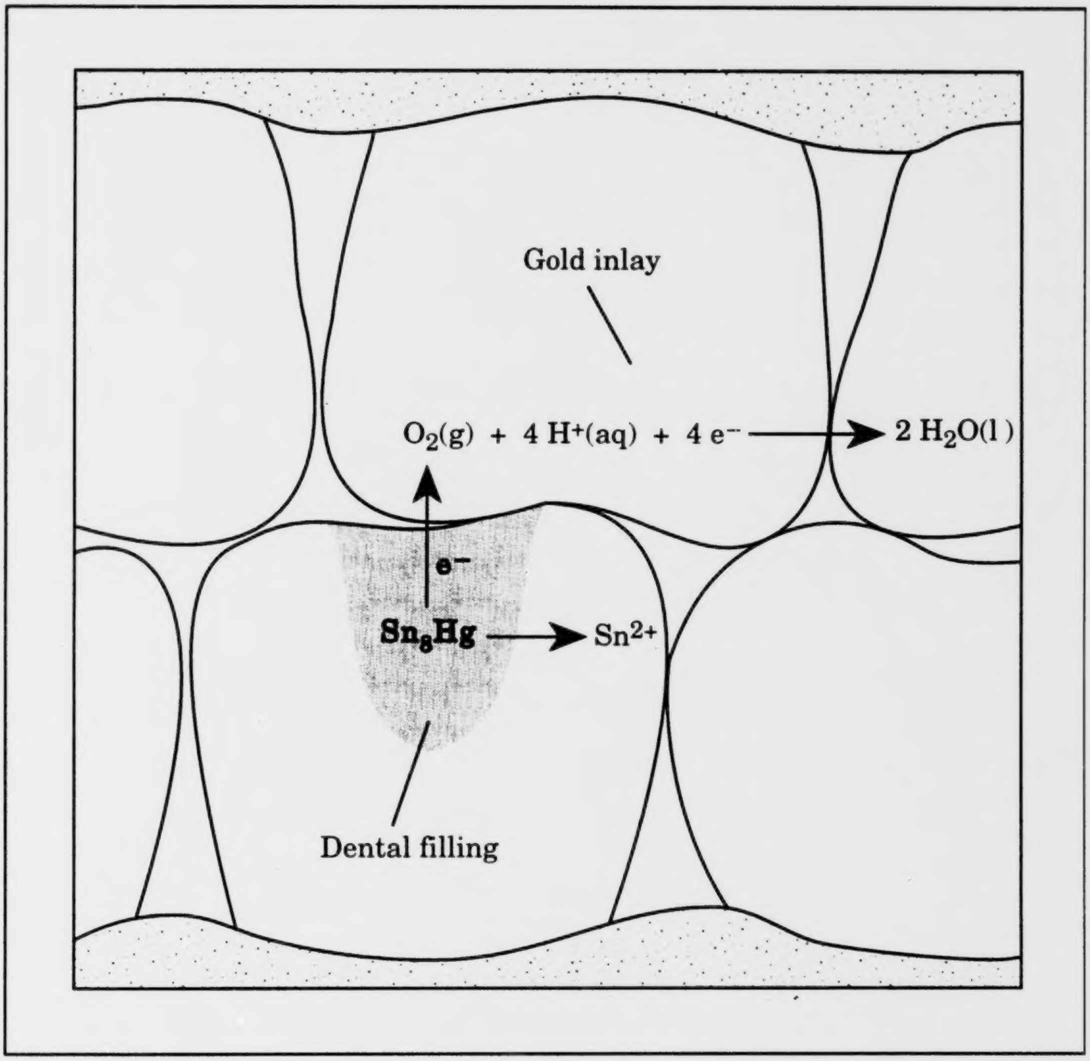

Dilute H_2SO_4 solution
Anode
Cathode
Battery


Battery
e⁻
e⁻
Anode
Cathode
Dilute H_2SO_4 solution

Oxidation: $2\ H_2O(l) \rightarrow O_2\,(g) + 4\ H^+(aq) + 4\ e^-$

Reduction: $4\ H^+(aq) + 4\ e^- \rightarrow 2\ H_2\,(g)$

Electrochemistry and Dental Fillings


Gold inlay
$O_2(g) + 4 H^+(aq) + 4 e^- \longrightarrow 2 H_2O(l)$
e^-
$Sn_8Hg \longrightarrow Sn^{2+}$
Dental filling

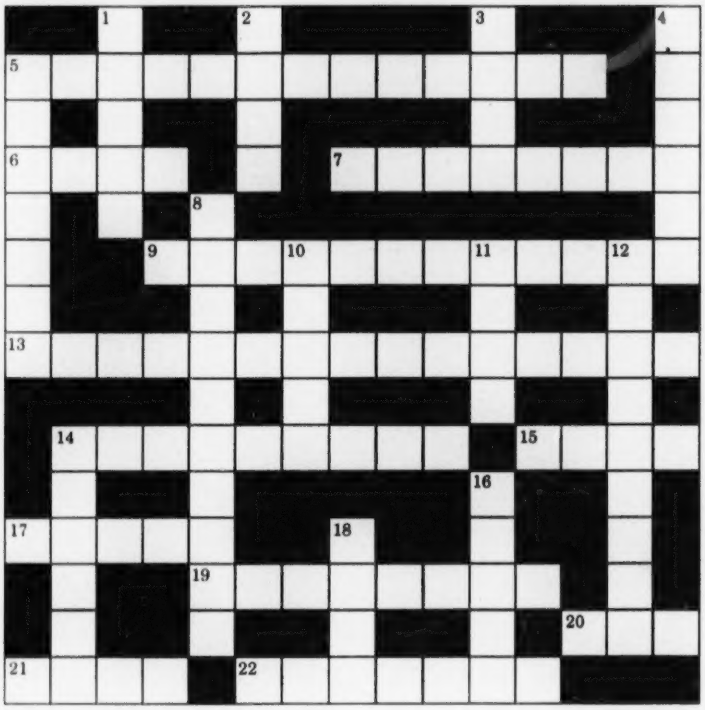
Word Search

A	S	E	G	D	I	R	B	T	L	A	S	N	P	E	Z	T
C	U	B	Z	N	J	T	F	O	V	F	T	N	D	D	J	C
O	H	A	V	O	L	T	K	L	G	O	J	O	B	C	S	Z
R	S	T	S	I	Q	V	V	Q	T	Y	H	I	G	U	G	M
R	N	L	I	N	E	F	Q	W	U	T	Y	T	U	Y	O	G
O	F	U	J	A	N	B	S	C	A	I	T	C	Q	L	W	M
S	N	G	J	B	G	A	I	C	S	Q	U	A	T	O	K	X
I	D	F	O	M	H	N	Y	C	Z	R	D	E	V	E	L	S
O	G	R	J	W	A	F	G	B	R	O	N	R	I	Q	K	Z
N	S	S	H	V	R	H	N	E	K	J	O	F	R	N	K	F
A	Q	W	L	A	I	T	N	E	T	O	P	L	L	E	C	R
L	P	A	R	W	S	T	I	Z	F	N	W	A	E	H	R	A
Y	G	L	Z	L	F	X	P	H	V	E	L	H	H	V	O	M

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Driving force in a voltaic cell. (2 words)
2. Either of two parts of an oxidation-reduction reaction. (2 words)
3. Electrical quantity measured in amperes.
4. Unit of electrical potential.
5. Electrode at which reduction occurs.
6. Negative ion.
7. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction
8. Process by which metals are oxidized in the atmosphere
9. State in which NaCl must be in order to conduct electricity
10. Device containing an electrolyte that connects the two compartments of a voltaic cell. (2 words)

Electrochemistry Crossword Puzzle



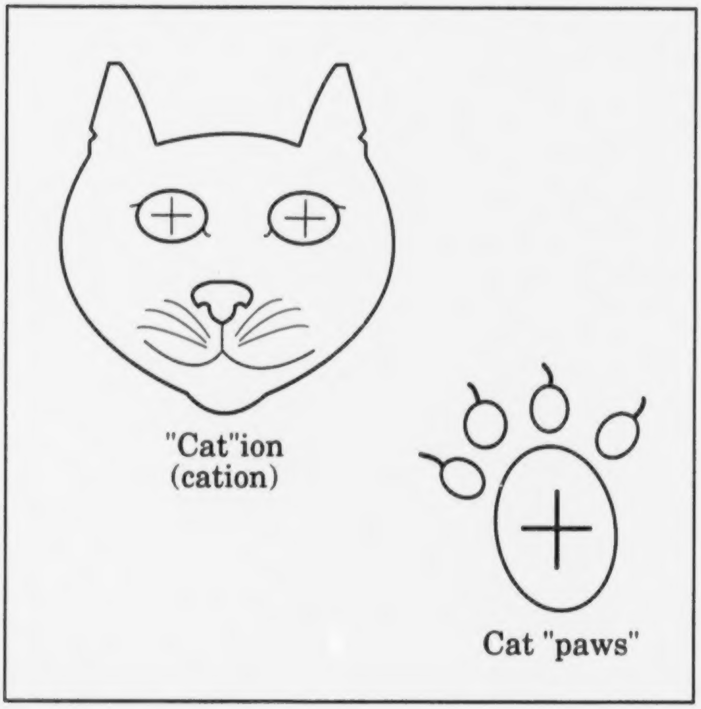
Across

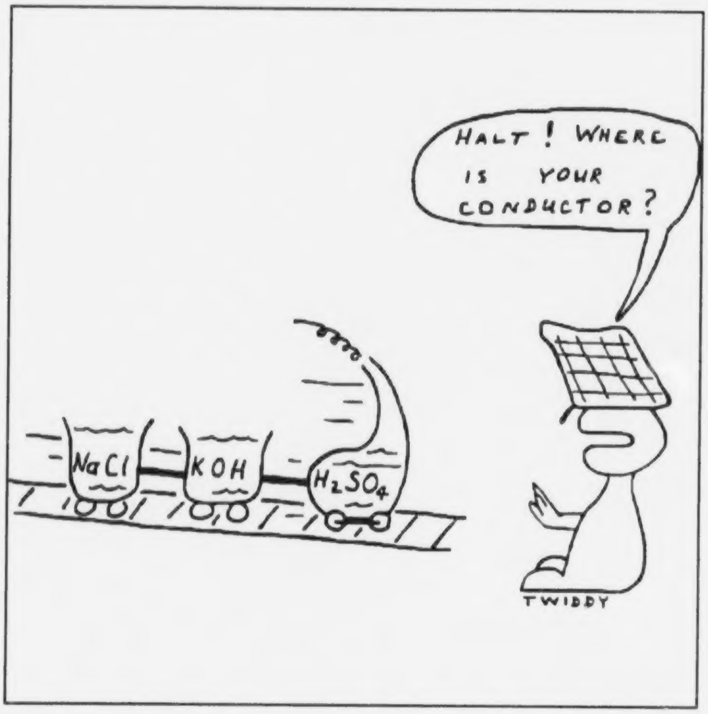
5. Driving force in a voltaic cell (2 words).
6. Fundamental quantity that must be measured in order to determine the rate of electron flow.
7. Substance added to fuel cells as fuel.
9. Either of two parts of an oxidation-reduction reaction (2 words).
13. Method of separating mixtures that contain charged molecular species.
14. Process by which metals are oxidized in the atmosphere.
15. Form of energy generated when electric current flows.
17. Word with ring denoting a support for laboratory devices.
19. Type of cell where chemical energy is converted to electrical energy spontaneously by a redox reaction.
20. Number of moles of electrons transferred when a mole of iron(II) is oxidized.
21. General term for a joule or coulomb, *e.g.*
22. Electrical quantity measured in amperes.

Down

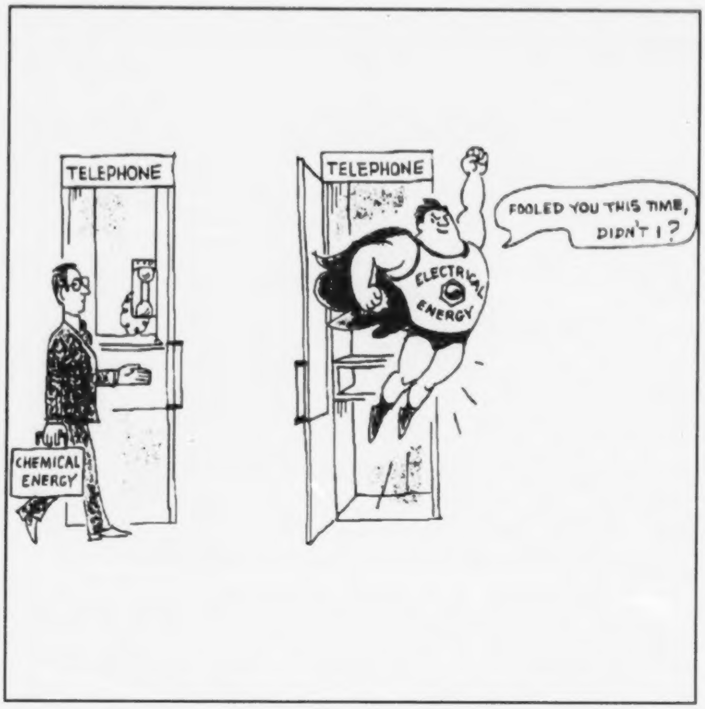
1. Laboratory holder for a test-tube.
2. Unit of electrical potential.
3. Element used in a Daniell Cell.
4. State in which pure NaCl must be in order to conduct electricity.
5. Electrode at which reduction occurs.
8. Device containing an electrolyte that connects the two compartments of a voltaic cell (2 words).
10. What an ion does when a voltaic cell discharges.
11. Center of a transformer or an apple.
12. Process whereby electrons are lost.
14. Positive ion.
16. Negative ion.
18. Word used to precede voltage to denote a condition where greater potential than calculated is needed to make a reaction proceed.

Across		Down	
5	CELL POTENTIAL	1	CLAMP
6	TIME	2	VOLT
7	REACTANT	3	ZINC
9	HALF REACTION	4	MOLTEN
13	ELECTROPHORESIS	5	CATHODE
14	CORROSION	8	SALT BRIDGE
15	HEAT	10	FLOWS
17	STAND	11	CORE
19	GALVANIC	12	OXIDATION
20	ONE	14	CATION
21	UNIT	16	ANION
22	CURRENT	18	OVER


"Cat"ion
(cation)
Cat "paws"


HALT! WHERE IS YOUR CONDUCTOR?
NaCl
KOH
H_2SO_4
TWIDDY

CHEM 13 NEWS, April 1984, p. 7. Reprinted with permission.


TELEPHONE
CHEMICAL ENERGY
TELEPHONE
ELECTRICAL ENERGY
FOOLED YOU THIS TIME, DIDN'T I?

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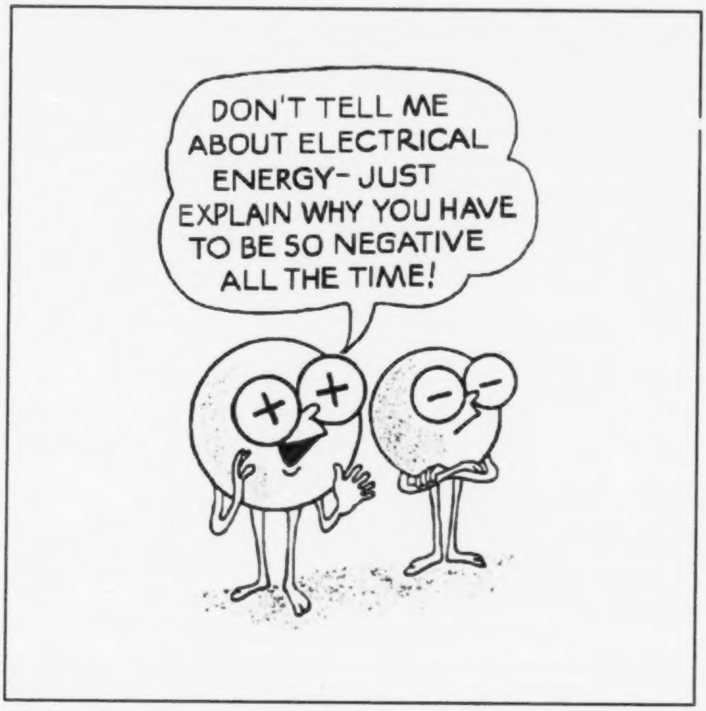
DON'T TELL ME ABOUT ELECTRICAL ENERGY- JUST EXPLAIN WHY YOU HAVE TO BE SO NEGATIVE ALL THE TIME!
+ +
− −

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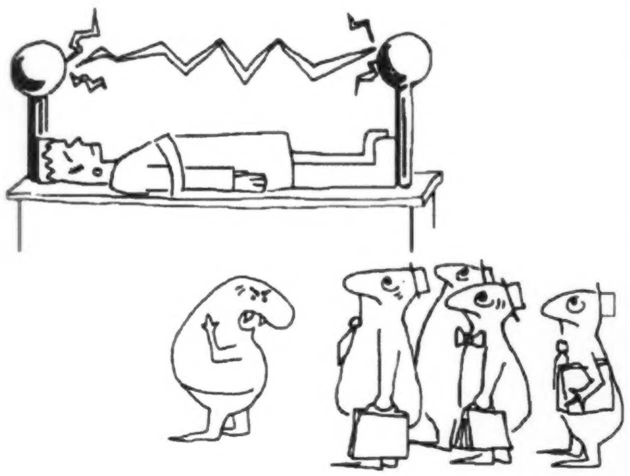
LONG
TERM
BATTERY

"Charged with a salt and battery."

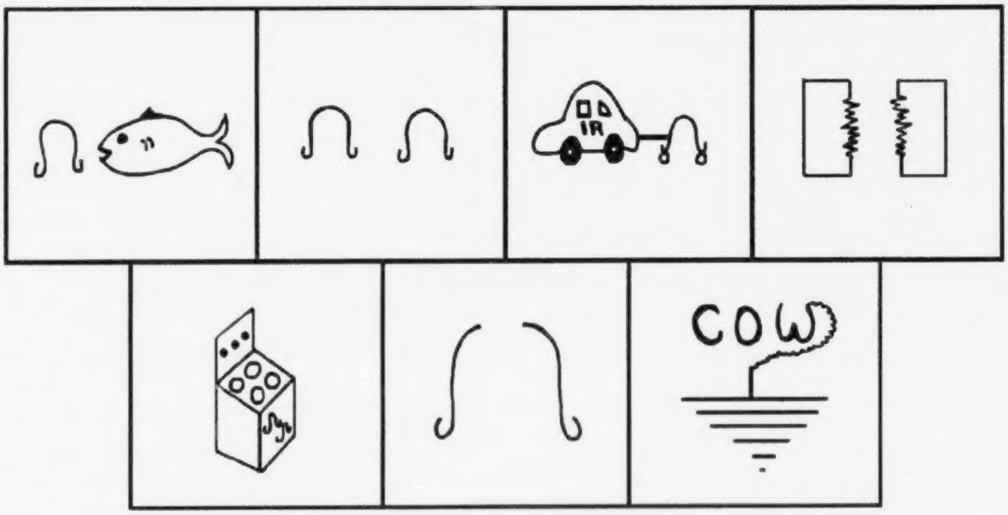
CHEM 13 NEWS, September 1978, p. 32. Reprinted with permission.



"Will you salesmen get off my back? I already HAVE a pair of electrodes."

CHEM 13 NEWS, March 1978, p. 1320. Reprinted with permission.

OH ME, OH MY, WHAT ARE THESE?



CHEM 13 NEWS, November 1980, p. 2. Reprinted with permission.

The answers to the quiz are:

1. ohm eater
2. ohm away from ohm
3. voltswagen pulling a mobile ohm
4. a pair of shorts
5. ohm, ohm on the range
6. product of a broken ohm
7. ground beef

A SourceBook Module

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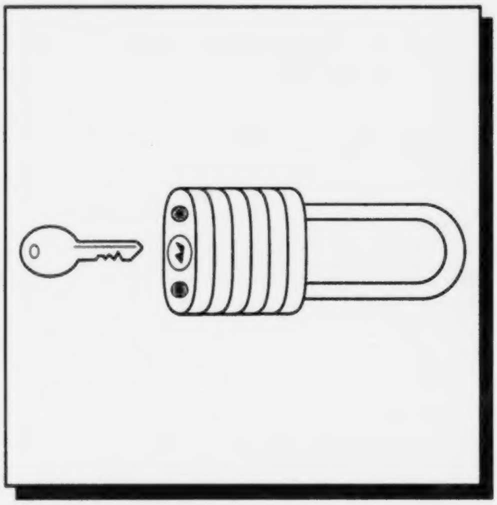
ENZYMES: BIOCHEMICAL CATALYSTS



ChemSource

Instructional Resources for Preservice and Inservice Chemistry Teachers

Topic Overview

CONTENT IN A NUTSHELL

In order to develop, grow, and reproduce, biological organisms depend upon thousands of chemical reactions. These reactions are called *biochemical reactions* and they occur because of specialized catalysts called enzymes. Although these enzymes do not force a reaction to occur that would not occur otherwise, they greatly increase the rates of body reactions that allow us to produce the energy we need to survive. If it were not for these biochemical catalysts, our body reactions would be so slow that life would be impossible.

Over 2,000 different enzymes in the body are known. This large number is necessary because enzymes are very particular about the reactions they catalyze; in fact, an enzyme may catalyze only one specific reaction. Enzymes are present in every body cell. One enzyme, which we will study in this module, is the enzyme *catalase*. It is produced in small organelles called *peroxisomes*, or *microbodies*. The enzyme catalase breaks down harmful hydrogen peroxide when it is produced excessively by body cells.

All enzymes are protein molecules. Although very large protein molecules, there is one primary area of the enzyme molecule where it chemically reacts with a substance called a **substrate**. This part of the molecule is called the **active site**. We can understand how an enzyme works by comparing its action on a substrate molecule to that of a key in a lock. The ridges and grooves of the key represent the active site of the enzyme, and, just as each key fits only a particular lock, each enzyme fits only a particular substrate. Also, just as the key opens the lock and is not destroyed in the process, enzymes—like other catalysts—emerge intact from a biochemical reaction.

Sometimes other molecules block enzymes or otherwise interfere with their reaction at the active site. These species include certain metal ions, insecticides, poisons, and bacterial toxins (like botulin that produces botulism). This process is called **enzyme inhibition**.

Enzymes act on substrates at a very rapid rate. One *catalase* enzyme molecule, for example, will completely break down 5.6 *million* hydrogen peroxide molecules *per minute*. Catalase is considered a relatively slow enzyme! One of the fastest enzymes, carbonic anhydrase, will break down 36 million carbonic acid molecules per minute. Now that's fast!

Since enzymes are protein molecules, they have the properties of proteins. They are denatured (rendered inactive) by high temperatures and many by extremes in pH. As the temperature of an enzyme-catalyzed reaction increases, the rate increases until the temperature is high enough to denature the protein.

PLACE IN THE CURRICULUM

This is an enrichment topic. It is most appropriate for the student who has had a general chemistry course.

CENTRAL CONCEPTS

1. Enzymes are special molecules that catalyze biochemical reactions.
2. All enzymes are protein molecules, and have general properties of proteins, *i.e.*, they are altered by pH and temperature changes.

3. Usually an enzyme acts only on a specific molecule (substrate), by first forming an enzyme-substrate complex, then breaking down into products and regenerating the enzyme.
4. Enzymes must be activated before they have the ability to function as a catalyst.
5. Only a small part of the enzyme molecule acts on the substrate. This part of the enzyme is referred to as the **active site**.
6. Enzymes can be inhibited when other substances block the active site.

RELATED CONCEPTS

1. pH
2. Chemical kinetics, especially the effects of temperature on reaction rates (see *Reaction Rates* module)
3. Characteristics of proteins
4. Structure of proteins
5. Denaturation of proteins

RELATED SKILLS

1. Collect and plot data on a graph.
2. Interpret a graph.
3. Measure temperature.
4. Work with general laboratory glassware.
5. Set up apparatus to collect gas by displacement of water.

Concept/Skills Development

LABORATORY ACTIVITY: STUDENT VERSION

Activity 1: The Effects of Temperature and Inhibitor on the Enzyme Catalase Extracted from Potato

Introduction

Catalase is an enzyme found in fairly high concentration in several fruits and vegetables. Its specific action is to decompose hydrogen peroxide into water and oxygen gas. This enzyme is also found in animal blood, and accounts for the "fizzing" when hydrogen peroxide is placed on an open wound. When the temperature is increased, the activity of enzymes also increases. However, a temperature is eventually reached where the protein enzyme is denatured and is no longer active. Enzymes can be inhibited by many substances. Catalase is inhibited by the copper(II) ion.

Purpose

To show that an enzyme can be isolated from a simple source, a potato, and that the activity of the enzyme depends upon temperature. To show that copper(II) ion is an inhibitor for the enzyme catalase.

Safety

1. Wear protective goggles throughout the laboratory activity.
2. Observe general safety rules when conducting this activity. None of the materials or solutions pose unusual safety hazards or require special handling.

NOTE: There are two alternate procedures for this activity. Your teacher will tell you which one to do.

Procedure A: Measuring Height of Foam Produced by Catalase

Procedure

1. Place potato pulp in a 50-mL graduated cylinder to a depth of about 3 cm.
2. Add 10 mL 3% hydrogen peroxide. Stir the cylinder quickly to mix the pulp and the hydrogen peroxide.
3. Note the time that a foam begins to form.
4. After 1 min, measure the height of the foam (from the top of the liquid to the top of the foam)
5. Measure and record the foam height each minute for at least 5 min.
6. Construct a graph that will allow you to record foam height on the y-axis and time on the x-axis. Label this *Graph 1.*

Effect of Temperature on Foam Height

1. Add potato pulp to a depth of about 3 cm to a second graduated cylinder and place the cylinder in a water bath for 2 to 3 min. *Your teacher will specify the temperature for your water bath.* Record the temperature.
2. Add 10 mL 3% hydrogen peroxide to another graduated cylinder.
3. Place both cylinders in a water bath for 5 min. *Your teacher will specify the temperature for your water bath.*
4. Remove both cylinders, quickly pour the 3% hydrogen peroxide into the potato pulp. Put the cylinder back into the water bath. Measure the foam height after 5 min and record your data.

Effect of an Inhibitor on Foam Height

1. Add potato pulp to a depth of about 3 cm to a fourth graduated cylinder. Add 10 drops copper(II) sulfate solution to the cylinder and stir.
2. Place 10 mL 3% hydrogen peroxide in a separate graduated cylinder.
3. Place both cylinders in a water bath at 20 °C for about 5 min.
4. Remove the cylinders from the water bath and add the 3% hydrogen peroxide to the potato pulp mixture.
5. Measure the foam height after 5 min. How does the foam height compare to that produced above without the copper(II) sulfate added?
6. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

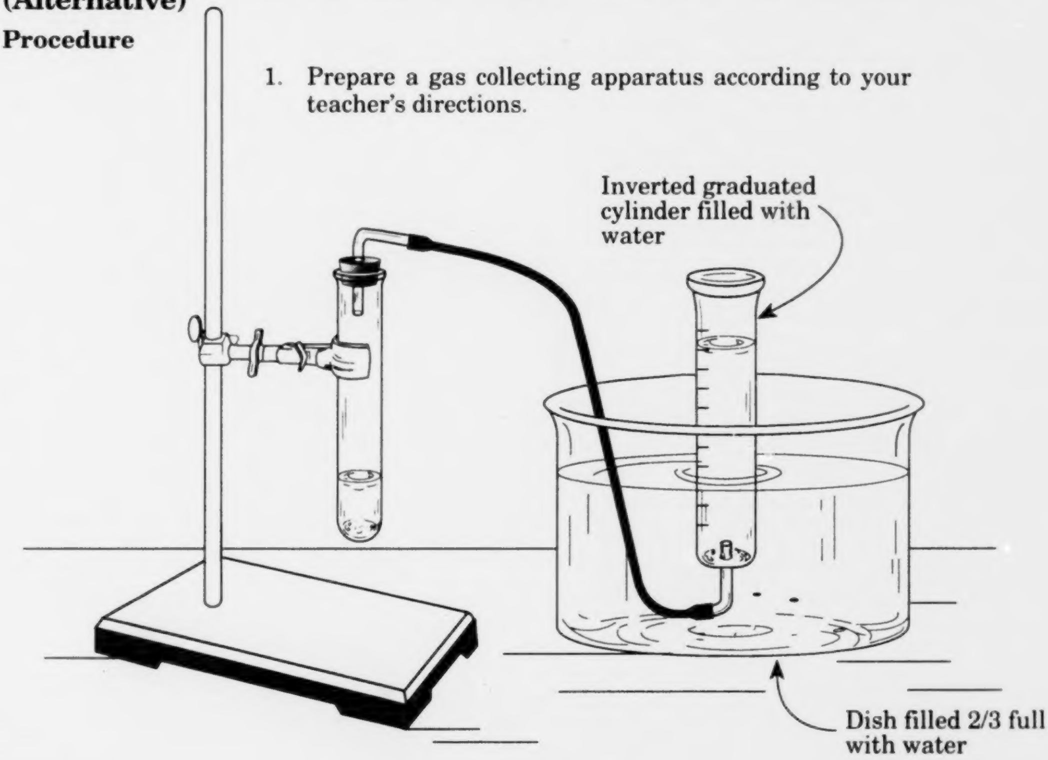
1. According to the data you plotted on *Graph 1*, how does the quantity of foam produced by the breakdown of hydrogen peroxide vary with time?
2. Your teacher will ask you to plot your data for temperature *vs.* height of foam on a master graph along with the data from your classmates.

   According to data plotted on the master graph, what is the effect of temperature on the rate of enzymatic action? At some point, did the quantity of foam *not* increase as temperature increased? How can you account for this?
3. According to your data, what effect did adding an inhibitor [copper(II) ions] have on the rate of foam production? Did the production of foam cease completely? How can you account for this?

Procedure B: Measuring Volume of Oxygen Produced by Catalase (Alternative)

Procedure

1. Prepare a gas collecting apparatus according to your teacher's directions.


Inverted graduated cylinder filled with water
Dish filled 2/3 full with water

2. Add about 10 mL 3% hydrogen peroxide to a small test-tube.
3. Quickly add 5 mL catalase extract (from the potato pulp).
4. Stopper the test-tube, swirl it once to mix the contents, and place the tubing leading from the stopper beneath an inverted, water-filled graduated cylinder. Record the time.
5. Record the volume of gas produced each minute for 5 min. Record your data.
6. Construct a graph that will allow you to record volume of gas produced on the y-axis and time on the x-axis. Label this *Graph 1*.

Effect of Temperature on Volume of Gas Produced

1. Add 10 mL catalase extract to a small test-tube.
2. Add 10 mL 3% hydrogen peroxide to a second small test-tube.
3. Place both test-tubes in a water bath for 5 min. *Your teacher will specify the temperature for your water bath.*
4. Remove both tubes, quickly pour 3% hydrogen peroxide into the catalase, stopper, and collect the gas produced as in the procedure above.
5. After 5 min, measure the amount of gas produced and record your data.

Effect of an Inhibitor on the Volume of Gas Produced

1. Place 10 mL catalase extract in a small test-tube.
2. Add 2 mL copper(II) sulfate solution to the tube.
3. Place 10 mL 3% hydrogen peroxide in a separate tube.
4. Place both tubes in a water bath at 20 °C for about 5 min.
5. Remove the tubes, add the 3% hydrogen peroxide to the catalase, shake to mix.
6. Quickly stopper the tube and collect the gas as you did above.
7. Measure the volume of gas produced after 5 min and record your data.
8. Thoroughly wash your hands before leaving the laboratory.

Data Analysis

1. According to the data you plotted on *Graph 1*, how does the volume of gas produced by the breakdown of hydrogen peroxide vary with time?
2. Your teacher will ask you to plot your data for temperature *vs.* volume of gas on a master graph along with the data from your classmates.

   According to the data plotted on the master graph, what is the effect of temperature on the rate of enzymatic action? At some point, did the volume of gas *not* increase as temperature increased? How can you account for this?
3. According to your data, what effect did adding an inhibitor [copper(II) ions] have on the rate of gas production? Did the production of gas cease completely? How can you account for this?

Implications and Applications

1. Enzymes have maximum activity at a temperature that depends upon the environment in which the organism thrives. Almost all enzymes found in the human body have maximum activity at 37 °C (98.6 °F). Can you explain this?
2. Most enzymes are inactivated at temperatures above 60 °C. What does this tell you about the effect of heat on body functions?
3. Increasing the temperature will increase the rate of reaction. In fact, it has been estimated that increasing the temperature by 10 °C will double the rate of enzymatic reactions. Do your data support this assumption?
4. Enzymes are typically found and used around the house. For example, meat tenderizer contains a *proteinase* that helps break down fiber and makes meat more tender. Would you add meat tenderizer to a steak while it is on the grill cooking? Some detergents contain a proteinase enzyme to help remove grime and stains, which are largely protein. Do you get best results from these detergents in *cold* or *warm* water?
5. Just as the copper ion inhibited the action of the enzyme catalase, other metal ions inhibit other specific enzymes. For example, mercury, lead, beryllium, and arsenic inhibit specific enzymes required for energy production. It is primarily for this reason that these metals are considered *toxic* and must be used with caution.

LABORATORY ACTIVITY: TEACHER NOTES

Activity 1: The Effects of Temperature and Inhibitor on the Enzyme Catalase Extracted from Potato

Major Chemical Concept

Enzymes are proteins that act as catalysts for biochemical reactions. The rate at which the catalyzed reaction occurs varies with temperature until the temperature is reached where the protein catalyst is deactivated. Catalysts are also inhibited by certain substances, including certain metal ions.

Level

This material is appropriate for first year or advanced high school chemistry classes.

Time

40-45 min

Expected Student Background

Students should understand the role of catalysts in a chemical reaction, the nature of proteins, and protein structure. They must be able to collect data, prepare a graph, interpret the graph, and make predictions. If students use *Procedure B*, they should know how to collect oxygen gas when it is produced in a chemical reaction.

Safety

Read the *Safety Considerations* in the *Student Version*. Materials used in these activities pose few safety problems. Students should not eat the potatoes used in the activity. Care should be taken with the hydrogen peroxide and copper(II) sulfate solutions. Wash with water if any gets on the skin.

Materials (For 24 students working in pairs)

Potatoes (red-skinned or Irish potatoes work well; see *Advance Preparation*)
0.1 M Copper(II) sulfate solution, $CuSO_4$ (see *Advance Preparation*)
Fresh 3% hydrogen peroxide, H_2O_2 (from drug or grocery store), 600 mL
60 Graduated cylinders, 50-mL
60 10- x 150-mm test tubes with stoppers
12 Metric rulers
4 Water baths (large beakers may be used)
Clock with second hand.
12 Graduated cylinders, 10-mL
12 Thermometers
4 Hot plates or burners
12 Gas collecting materials (for *Procedure B*): one-hole stopper (to fit test-tube) with glass tubing (or plastic tubing) through hole; flexible hose attached to the glass tube, leading to water trough filled 2/3 with water and water-filled inverted
12 Graduated cylinders, 100-mL

Advance Preparation

0.1 M Copper(II) sulfate solution: Dissolve 2.5 g copper(II) sulfate pentahydrate ($CuSO_4 \cdot 5H_2O$) in 100 mL water.

Potato extract. Prepare the pulp and extract as follows: Peel 2-3 Irish potatoes. Cut the potatoes into pieces and place them in a blender or food processor. Add about 50 mL water and thoroughly grind the potatoes. Let the mixture stand for a few minutes, and filter it through cheesecloth to obtain a pulp-free solution. The enzyme catalase is soluble, so it is in both the pulp and the liquid extract.

Drug store variety hydrogen peroxide is 3%. *It must be fresh!*

Pre-Laboratory Discussion

1. Review graphing.
2. Discuss general properties of proteins.
3. Discuss general principles of reaction kinetics (see *Reaction Rates* module).
4. Assign *Procedure A* or *Procedure B* (or both).
5. Assign temperature to groups. Recommended temperatures are:

   Approximately 0 °C (ice bath)
   Room Temperature, 20-25 °C
   Body Temperature, 37-40 °C
   High Temperature, 70-80 °C
6. Explain the nature of a *foam* (a substance that has trapped a gas).

Teacher-Student Interaction

The success of this activity requires a *group* effort. Students must be encouraged to contribute their data, even if it does not seem to *fit* a pattern on the group graph. It is imperative that you move from group to group to be sure that they are doing the activities correctly and safely. Question students about their results as they are conducting the activities.

Anticipated Student Results and Answers to Data Analysis

1. Volume of gas (foam height) *vs.* time: The volume of gas (foam height) will increase regularly, but may tend to level off after 5 min or so.
2. Temperature *vs.* volume of gas (foam height): The volume of gas (foam height) will increase as the temperature increases, but will drop off at higher temperatures.
3. Inhibitor and enzyme: The volume of gas (foam height) will be much less, but may not drop to zero.

Answers to Implications and Applications

1. This is normal body temperature. As cells evolved, they adapted to the temperature that would produce maximum product from enzyme action.
2. High body temperature can inactivate protein, including the protein component of enzyme systems.
3. Yes. Class data will generally show a doubling of reaction rate with each 10 °C increase in temperature.
4. You would not add meat tenderizer to meat while it was cooking, since the thermal energy would inactivate the enzymes in the meat tenderizer. The best results would be in warm water. This would increase the activity of the enzymes. If the water were too hot, however, the enzymes would become inactivated.

Post-Laboratory Discussion

1. If data seem to be "nonfitting" to graph, discuss experimental sources of error. Ask students to suggest possible sources of error.
2. Discuss the general relationship between rates of reactions and temperature.
3. Discuss the evidence for enzyme inhibition and gas production.
4. Ask students to suggest applications.

Possible Extension

1. Try other simple enzymes. Refer to college biochemistry texts for examples.
2. pH also affects enzymes. Design studies to show the effect of pH on catalase and other enzymes.
3. Have students perform library research and write reports pertaining to enzymes.

Assessing Laboratory Learning

Although you should not penalize students for not obtaining data that "fit" the graphs, they should be evaluated on their laboratory technique, ability to work as a group, and ability to record data.

References

These activities are modified from those found in:

Borgford, C., and Summerlin, L. (1988). *Chemical activities.* Washington: American Chemical Society.

Summerlin, L., Borgford, C., and Ealy, J. (1987). *Chemical demonstrations: A sourcebook for teachers, Volume 2.* Washington: American Chemical Society.

DEMONSTRATIONS

CAUTION: Use appropriate safety guidelines in performing demonstrations.

Demonstration 1: Calcium ions, Rennin, and the Coagulation of Milk

Purpose

This demonstration illustrates the importance of a metal ion as an enzyme activator.

Safety

There are no special precautions for this demonstration. Students should be careful with the warm water.

Materials

Whole milk
Sodium citrate
Rennin (sold in grocery stores as Rennilase or Rennet), available from biological supply companies
Beaker of warm water

Procedure

Place 15-mL samples of fresh, whole milk in each of two large test-tubes. To one, add a pinch (match-head size) of sodium citrate, and stir. Add nothing to the other test-tube. Add a small amount of the enzyme to both test-tubes (about 5 drops of rennilase, or about 1/4 tablet of rennet). Place both test-tubes in a larger beaker of warm water for three to 5 min. Compare the results.

Remarks

Rennin is an enzyme that converts the soluble milk protein caseinogen into the insoluble protein paracasein, producing the curd that can be processed into cheese or other milk products. Calcium ion is required to activate the enzyme rennin. In this demonstration, calcium ion is removed by precipitating it as calcium citrate. In the tube containing the sodium citrate, calcium citrate forms as a precipitate:

$$Na_3C_6H_5O_7(aq) + Ca^{2+}(aq) \rightarrow NaCaC_6H_5O_7(s) + 2Na^+(aq)$$

In the absence of calcium ions, the rennin is not activated and cannot catalyze the reaction to convert caseinogen into paracasein; thus, the milk does not coagulate.

Reference

Borgford, C., and Summerlin, L. (1988). *Chemical activities.* Washington, DC: American Chemical Society.

Demonstration 2: Pectinase and Apple Juice Production

Purpose

The activity of a typical enzyme, pectinase, is illustrated in this demonstration. Pectinase breaks the cell wall of fruit, producing more "juice."

Safety

Students should not eat or drink any of the material from this demonstration.

Materials

Applesauce
Pectinase enzyme (available from chemical or biological supply companies)
Beakers
Droppers
Graduated cylinder
Filter paper

Procedure

Measure about 50 mL applesauce into each of two small beakers. Add 1-mL pectinase to the applesauce in one beaker; add nothing to the other beaker. Stir the applesauce in both beakers, and allow to stand for 10 min. Filter each, and measure the amount of juice produced from the applesauce with and without the enzyme.

Remarks

Pectin is a large polysaccharide that is present in fruit cells. It prevents material, including the juice of the fruit, from settling out in the cells. When this large molecule is acted on by the enzyme pectinase, the juice can be easily separated. The pectinase enzyme ruptures the cells of the apple, allowing the juice to settle and to be easily separated from the other material in the cell.

Extensions

Students should be encouraged to devise activities to answer the following questions:

1. What is the relationship between the amount of enzyme added and the amount of juice produced? *[The greater amounts of enzyme will lead to production of more juice.]*
2. What is the relationship between time of standing and the amount of juice? *[Typically more juice will be produced.]*
3. What is the relationship between the temperature and the amount of juice produced? *[Enzyme activity is dependent on the temperature.]*

Reference

Borgford, C., and Summerlin, L. (1988). *Chemical activities.* Washington, DC: American Chemical Society.

Demonstration 3: Catalase in Turnips, Potatoes, and Rutabagas

Purpose

To demonstrate the activity of a typical enzyme.

Safety

Be careful when slicing the vegetables in the demonstration.

Materials

3% Hydrogen peroxide (drug store variety)

Potatoes, turnips, rutabagas—fresh, whole; animal tissue, such as liver, works well

Procedure

Slice several potatoes, turnips, and rutabagas into 3-4 smaller pieces. Place each sample in a separate shallow dish. Pour sufficient 3% hydrogen peroxide to cover the exposed surface. Pass the dishes around to allow students to observe the bubbling on the cut surfaces of the vegetables. You might observe cut side *vs.* skin side of vegetables.

Remarks

Catalase, found naturally in the vegetables used here, reacts with the hydrogen peroxide to rapidly decompose it into water and oxygen gas.

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2\,(g)$$

Catalase is the enzyme that specifically decomposes hydrogen peroxide into water and oxygen gas. Perhaps you have noticed that when hydrogen peroxide is placed on an open wound, fizzing and bubbling occur. This reaction is caused by catalase in the blood decomposing the hydrogen peroxide and releasing oxygen gas.

Reference

Summerlin, L. (1979). *Chemistry of common substances.* Morristown, NJ: Silver-Burdett.

GROUP AND DISCUSSION ACTIVITIES

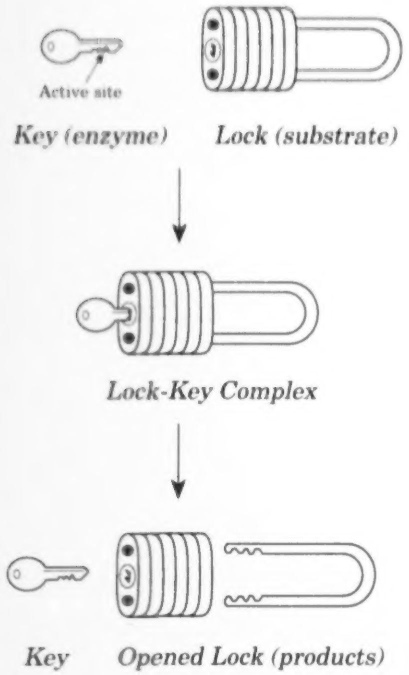
Counterintuitive Example

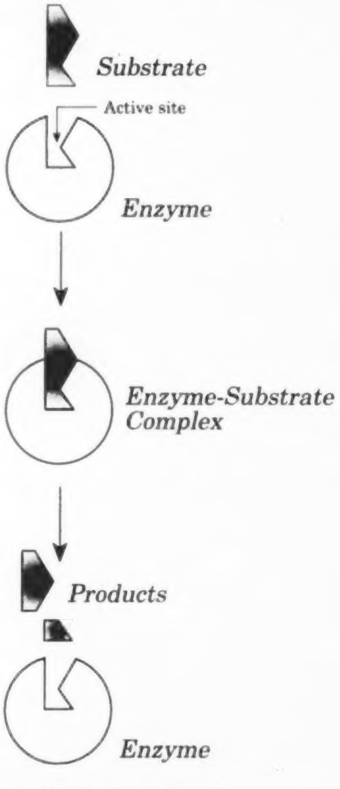
Proteins (as in meats) are broken down by enzymes produced in the pancreas. The pancreas is composed of protein. Why don't these enzymes digest the pancreas? *[Enzymes, including those produced by the pancreas, are produced as* inactive *enzymes, or* pro-enzymes. *When they get to the site where they are needed, they are then "activated" by another substance. The pancreas produces inactive trypsinogen, which is activated in the small intestine by the "activator" secreted by the cells in the wall of the small intestine. The "active" enzyme is* trypsin.*]*

Metaphors and Analogies

1. Enzymes are usually confined to the area where they are needed to catalyze specific reactions. These areas include the liver, pancreas, and muscle (including heart). When they are found in significant concentration outside those areas, it means that the tissue or organ containing those enzymes has been damaged, allowing them to "escape." This phenomenon can be compared to a prison, where felons are grouped together. If these convicts are seen outside the prison it is cause of great concern!
2. Enzymes bringing together substrate molecules to accelerate a reaction can be compared to a chaperon (enzyme) at a dance bringing bashful students (substrates) together and getting them to know each other (form products).

3. The mechanism of enzyme action is well illustrated by the "lock and key" analogy. The "lock" represents the substrate and the "key" represents the enzyme. Just as a key is specific for a certain lock, an enzyme is specific for a certain substrate. Also, just as the enzyme has an "active site" where the chemistry occurs when it combines with a substrate, the key has grooves and notches that must match those in the lock mechanism. After the key "opens" the lock remains intact, just as the enzyme reappears intact after acting on a substrate.

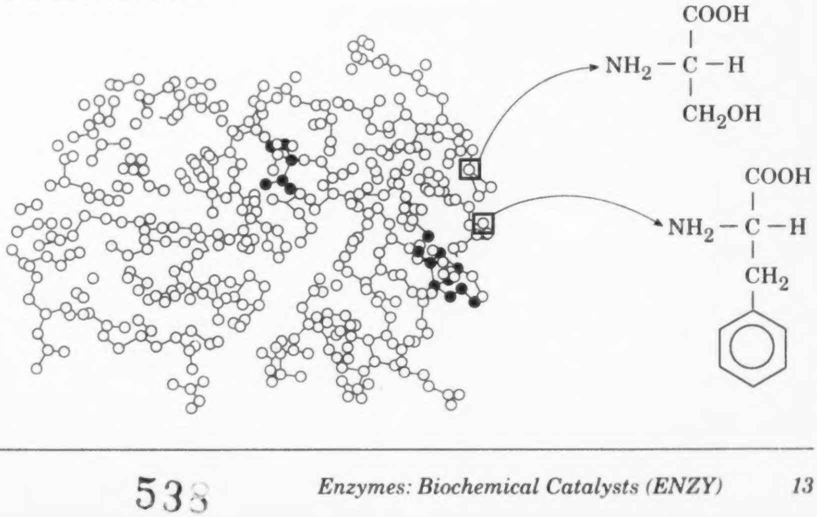

Active site
Key (enzyme)
Lock (substrate)
Lock-Key Complex
Key
Opened Lock (products)


Substrate
Active site
Enzyme
Enzyme-Substrate Complex
Products
Enzyme

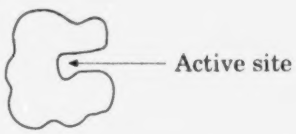
4. A new theory of enzyme activity is the "induced fit" model, where the enzyme conformation changes to accommodate the substrate. This is somewhat like a glove changing to fit the hand inserted into it.

Pictures in the Mind

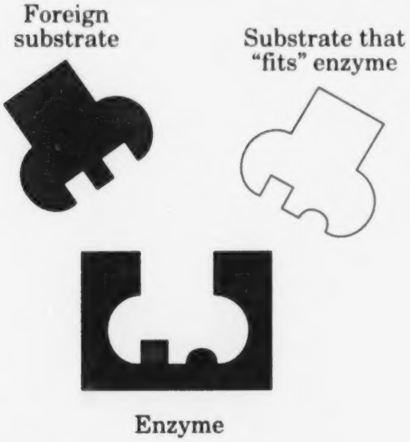
1. Draw a picture of an enzyme molecule. Each circle represents an *amino acid*. Typical amino acids are shown here.


COOH
NH2 — C — H
CH2OH
COOH
NH2 — C — H
CH2

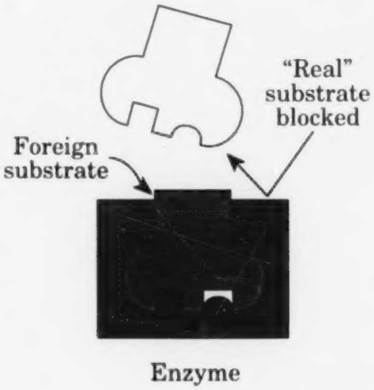
2. Draw a picture of an enzyme molecule showing its *active site*.
3. Draw a picture of a substrate molecule that "fits" the active site of an enzyme, and one that does not "fit."



FITS



DOES NOT FIT



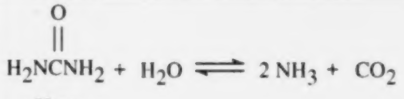
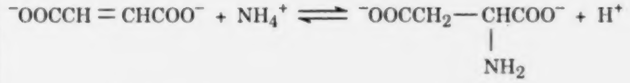
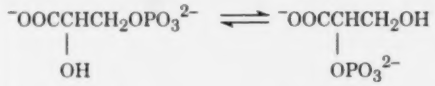
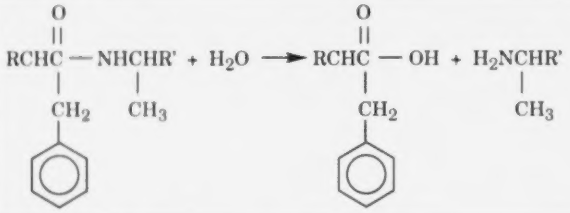
Pattern Recognition

1. Enzymes were previously named by adding the suffix *-ase* to the name of the substrate upon which they acted. For example:
   - *Sucrase* is the enzyme that acts to break down *sucrose*.
   - *Lipase* breaks down *lipids*.
   - *Peroxidase* breaks down *peroxides*.
2. Enzymes are now properly classified according to the *type* reaction they catalyze. The classes include:
   - *Oxido-Reductases* involve donating and accepting electrons.
   - *Hydrolases* add water to break chemical bonds. (Most digestive enzymes are in this category.)
   - *Transferase* transfers groups (amino, phosphate, carboxyl, *etc.*) from one molecule to another.
   - *Lyases* break (or form) double bonds.
   - *Isomerases* change one isomer into another.
   - *Ligases* allow C—C, C—N, and C—S bond formation.

Problem Solving/Decision Making

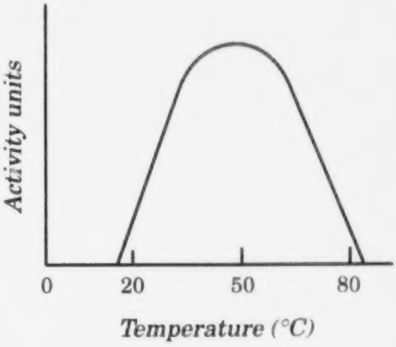
1. Humans cannot digest wood, paper, and other cellulose materials because cellulase, the enzyme required to break these materials down into usable glucose, is not present in the body. It is produced by bacteria that live in the intestinal tracts of termites, horses, goats, *etc.* Suppose that someone suddenly found a way for this enzyme to exist in humans. Reflect on the wide implications of such a possibility, and write a short paper describing what effect this might have on humans.

2. Each of the five enzymes listed below specifically catalyzes the reaction shown. Considering the classification scheme for types of enzymes given in the *Pattern Recognition* section, decide to which group each of the enzymes belongs:

Enzyme	Reaction
(a) urease	 $H_2N\overset{O}{\overset{\|}{C}}NH_2 + H_2O \rightleftharpoons 2\,NH_3 + CO_2$ Urea
(b) aspartase	 ${}^-OOCCH = CHCOO^- + NH_4^+ \rightleftharpoons {}^-OOCCH_2{-}\underset{NH_2}{\underset{\|}{C}}HCOO^- + H^+$ Fumarate L-Aspartate
(c) phosphoglyceromutase	 ${}^-OOC\underset{OH}{\underset{\|}{C}}HCH_2OPO_3^{2-} \rightleftharpoons {}^-OOC\underset{OPO_3^{2-}}{\underset{\|}{C}}HCH_2OH$ 3-Phosphoglycerate 2-Phosphoglycerate
(d) succinate dehydrogenase	${}^-OOCCH_2 - CH_2COO^- + FAD \rightleftharpoons {}^-OOCCH = CHCOO^- + FADH_2$ Succinate Coenzyme Fumarate Reduced coenzyme
(e) chymotrypsin	

3. What would happen to the rate of an enzyme catalyzed reaction if the concentration of the enzyme is doubled? What would happen to the rate if the concentration of the substrate is doubled? Explain your answers on the basis of kinetics (see *Reaction Rates* module).

4. The activity of a certain bacterial enzyme at various temperatures is shown in the following graph:



Suppose a person is infected with this bacterium.

a. What happens to the activity of this enzyme if the person has a high fever? *[It increases (approaching 50 °C optimum).]*

b. What happens to the activity of this enzyme if the person's temperature drops below 37 °C? *[It decreases.]*

c. At what temperature does this enzyme have its maximum activity? *[50 °C.]*

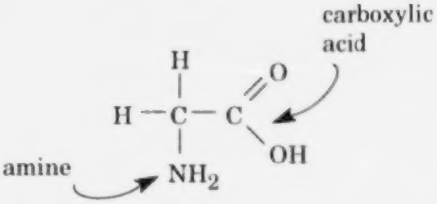
5. What is the relationship between the water-soluble vitamins and enzyme function? Find out which vitamins are related to which enzymes and which disease or deficiency state results when that vitamin is missing. You can research this problem in the library. Ask a physician or pharmacist for advice.

TIPS FOR THE TEACHER

Enzymes are *proteins*. The word *protein* is from the Greek *protos*, meaning "first," or "of prime importance." This very important group of biochemical molecules are, indeed, of prime importance. They are found in every living cell and they are extremely diverse in form and function. No other biochemical substance serves so many different functions as proteins. Proteins also provide the primary structural component of the cell and also direct the synthesis of cellular materials. Proteins are responsible for transporting materials in the body, including oxygen, ions, lipids, and other substances to the cells where they are needed for metabolism.

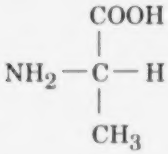
All proteins share common features. They are very large molecules with molar masses ranging from 5,000 to several million daltons. Such molecules are often called *macromolecules*. In addition to containing carbon, oxygen, and hydrogen, all proteins contain nitrogen and some contain sulfur. When proteins are consumed as nutrients, it is necessary to remove the nitrogen and excrete it. Some animals excrete nitrogen as ammonia, but ammonia is toxic to humans and nitrogen is excreted as the compound called *urea*. The most important feature of proteins is that they are all polymers of smaller units called *amino acids*.

Amino acids are compounds containing both an *amino* group and a *carboxylic acid* group. All amino acids found in proteins have these two groups on the same carbon atom.



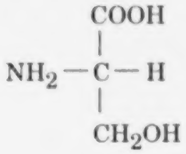
Glycine

Of 150 known amino acids, only about 20 commonly occur in human protein. Eight of these common amino acids cannot be synthesized by the body and must be included in the diet. These are called **essential amino acids**. All of these amino acids have the basic structure as shown, and differ only in the group attached to the carbon atom bearing the amino and carboxylic acid groups. Natural amino acids have the L-configuration. Three typical amino acids are:



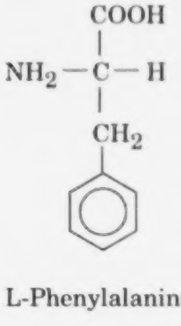
$$\begin{array}{c} \mathrm{COOH} \\ | \\ \mathrm{NH_2-C-H} \\ | \\ \mathrm{CH_3} \end{array}$$

L-Alanine



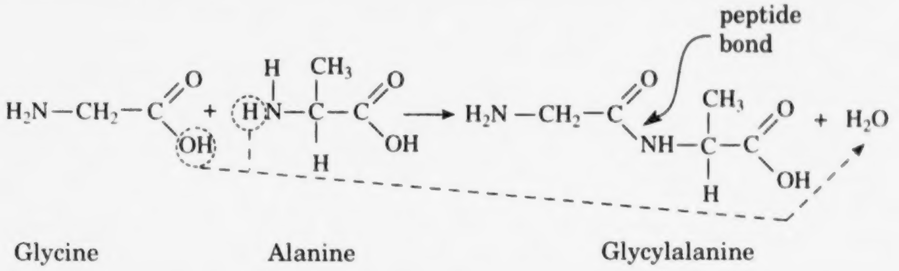
$$\begin{array}{c} \mathrm{COOH} \\ | \\ \mathrm{NH_2-C-H} \\ | \\ \mathrm{CH_2OH} \end{array}$$

L-Serine

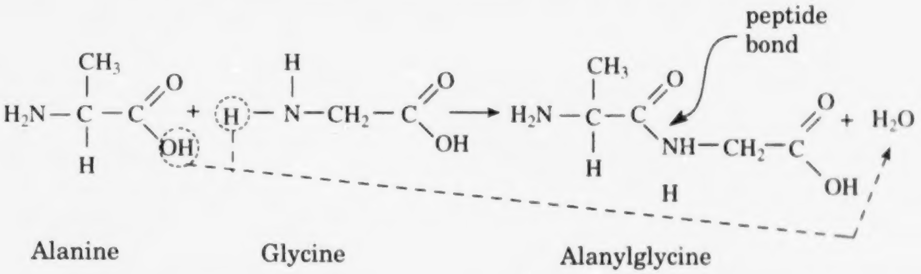


L-Phenylalanine

Amino acids can combine easily by removing a water molecule, joining the amino end of one amino acid with the carboxyl end of another. The resulting bond is called a **peptide bond**. The process of peptide bond formation is identical to the process of condensation polymerization forming nylon (see *Polymers* module). For example, the amino acids glycine and alanine can combine to form a dipeptide, glycylalanine:

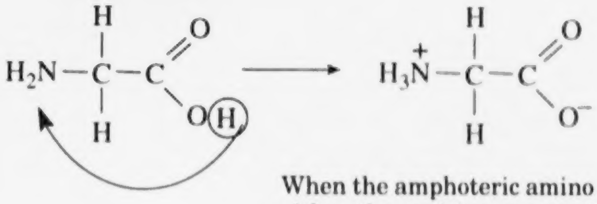


When several amino acids are joined in such a manner, the resulting structure is called a **peptide**. Examining the way the peptide bond is formed, it is obvious that it could form by connecting the amino and carboxyl end of either amino acid. Thus, there could be two different ways to join glycine and alanine. For example,



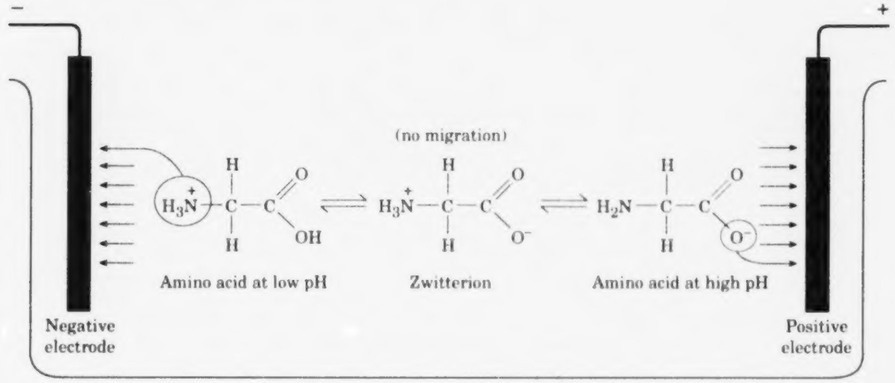
Although glycylalanine and alanylglycine, contain the same two amino acids, they are *two completely different compounds.* In fact, if 4 different amino acids are combined, and if each amino acid can appear more than one time in this sequence of four, it would be possible to form *256 different compounds.* How many different compounds can be formed from all 20 amino acids? A quick calculation shows that 10^{26} compounds are possible! This mind-boggling number clearly shows the possibility of billions of different proteins, all consisting of only 20 amino acids! The order in which these amino acids are strung together is the protein's **primary structure**. When proteins are digested the process is reversed and peptide bonds are broken to release amino acids. Since water is required, the process is called **hydrolysis**.

As polypeptides get larger, cross-linking occurs. An important cross-linking bond is the *disulfide* bond, which is formed by thiols (–SH groups). Thiols can react with heavy metals, such as mercury and lead, which results in the alteration of protein structure. This structural alteration accounts for heavy metal toxicity. As the protein molecule becomes more complex, other bonds—*ionic, hydrogen, and hydrophobic / hydrophilic*—become important in holding the giant molecule together. We can effectively break these bonds and denature or destroy the protein using such things as heat, acids and bases, alcohol, and detergents. Perhaps you can think of some practical examples of protein denaturation (*e.g.*, frying or whipping egg white).



Amino acids also have a unique property that allows separation. They have both acid and base properties. Such structures are said to be amphoteric, and the double-ion produced is called a **zwitterion**.

When the amphoteric amino acid is placed in an electric field, it tends to migrate to either the positive or negative pole, depending on the strength of its positive or negative charge. However, if the pH of the solution containing the amino acid is adjusted to the point that the true zwitterion exists, the amino acid has no net charge and it will not migrate in an electric field. This point is called the isoelectric point for the amino acid, and it is a characteristic property of each individual amino acid—much like a fingerprint is characteristic of each individual. The procedure just described is called electrophoresis, and provides a simple way to identify proteins (see *Forensic Chemistry* module).



Scientists now know a great deal about proteins, and are constantly learning more about these complex and fascinating macromolecules.

Language of Chemistry

Proteins

α-helix model of protein structure in which intrachain hydrogen bonding holds polypeptide coils together in a helix.

amino acid compound containing both an amino group and a carboxylic acid group in the same molecule. Amino acids are the building blocks of protein.

amphoteric describing a substance that can act as either an acid or a base.

denatured protein protein that does not exist in the native state, but has been altered.

disulfide bonds bonds formed between sulfur atoms of polypeptide chains.

electrophoresis technique of regulating the migration of amino acids in an electric field through pH manipulation.

isoelectric point specific pH at which an amino acid has no net charge and exists as a zwitterion.

peptide bond bond formed between the carboxylic acid group of one amino acid and the amino group of another amino acid.

primary protein structure sequence of amino acids in a polypeptide.

protein polymer linked together by peptide bonds that gives amino acids upon hydrolysis. Proteins are found in plant and animal tissue.

secondary protein structure orientation of a peptide chain in space.

zwitterion molecule bearing both a positive and a negative charge. Neutral protein molecules usually exist as zwitterions where the acidic carboxylate hydrogen ($-COOH$) is transferred to the amino group ($-NH_2$) resulting in the presence of $-COO^-$ and $-NH_3^+$ groups in the protein.

Enzymes

apoenzyme inactive, protein component of a holoenzyme.

coenzyme activator that converts an apoenzyme into an active enzyme.

enzyme biocatalyst protein produced by living cells; enzymes regulate cellular reactions without themselves being altered or destroyed.

holoenzyme complete, activated enzyme.

metalloenzyme holoenzyme (active enzyme) produced by an apoenzyme and a metal-ion activator.

molecular rate number of substrate molecules acted upon by a single enzyme molecule per minute.

phenylketonuria build-up of phenylpyruvate in the body due to the absence of the enzyme phenylalanine hydroxylase.

zymogen (also often called **proenzyme**) inactive form of an enzyme.

Common Student Misconceptions

1. **"Enzymes *cause* a reaction to occur."**

   Enzymes (or other catalysts) will not cause a reaction to "go" if that reaction won't "go." However, the enzyme does lower the potential energy barrier (activation energy) and permits the reaction to proceed at a much more rapid rate.

2. **"Enzymes (as found in dietary supplements, over-the-counter drugs, *etc.*) are effective when ingested."**

   Enzymes (and other proteins) are degraded in the stomach or intestine and are not absorbed. Only a few digestive enzymes are taken orally, and their effectiveness is questionable.

3. **"Only a few enzymes are needed in the body."**

   There are over 2,000 *known* biochemical reactions that require specific enzymes; thus there are over 2,000 different enzymes.

4. **"Enzymes are found only in animal cells."**

   Enzymes are found in both plants and animals, although they are all proteins. Turnips and potatoes, for example, contain the enzyme catalase that breaks down hydrogen peroxide into oxygen and water. The same enzyme is found in human blood.

5. **"Enzymes (catalysts in general) speed up a reaction without entering into the reaction."**

Although a catalyst, or an enzyme, is the same after a reaction has occurred and appears to be unchanged, it plays a very active part in the reaction and undergoes changes that involves breaking and forming new bonds in the enzyme as well as in the substrate. Just as with other catalysts, the enzyme forms an *enzyme-substrate complex*, equivalent to the activated complex formed by most chemical catalysts.

6. **"Enzymes increase the amount of product formed by changing the *equilibrium point* of a reaction."**

Enzymes do *not* change the equilibrium point of a chemical reaction. They only affect the rate at which the equilibrium point is approached.

HISTORY: ON THE HUMAN SIDE

1831 Berzelius assumed that certain substances contain a *"catalytic force"* that permitted them to accelerate a reaction.

1833 Jean Persoz and Anselme Payer prepared a malt extract that converted starch to dextrins. They called this substance *diastase* from the Greek word that means "breaking." We now call this enzyme amylase.

1836 Theodor Schwann showed that a living cell (yeast) was required for fermentation.

1845 Mialhe discovered diastase in saliva and called it *animal diastase.*

Bouchardat, Sandras and Valentin found the same substance in pancreatic juice and suggested that the pancreas be called the *abdominal salivary gland.*

1876 Kuhne suggested the name enzyme, from the Greek word meaning *leaven*. Such substances were previously called "ferments."

1883 As more enzymes were discovered, Duclaux suggested that all enzymes be named to end in *-ase*, as in *diastase.*

1897 Eduard Buchner accidentally isolated *zymase* from yeast and showed it to be the active enzyme in yeast.

Bertrand proposed the first enzyme-activation scheme: calcium and pectase to form the active enzyme, pectinase.

1913 Michaelis performed the first kinetic studies on enzymes.

1926 James Sumner crystallized *urease* from jack bean meal and showed that it was a protein.

1930 John Northrop crystallized *pepsin* and showed that it was also a protein. He suggested that *all* enzymes are protein.

HUMOR: ON THE FUN SIDE

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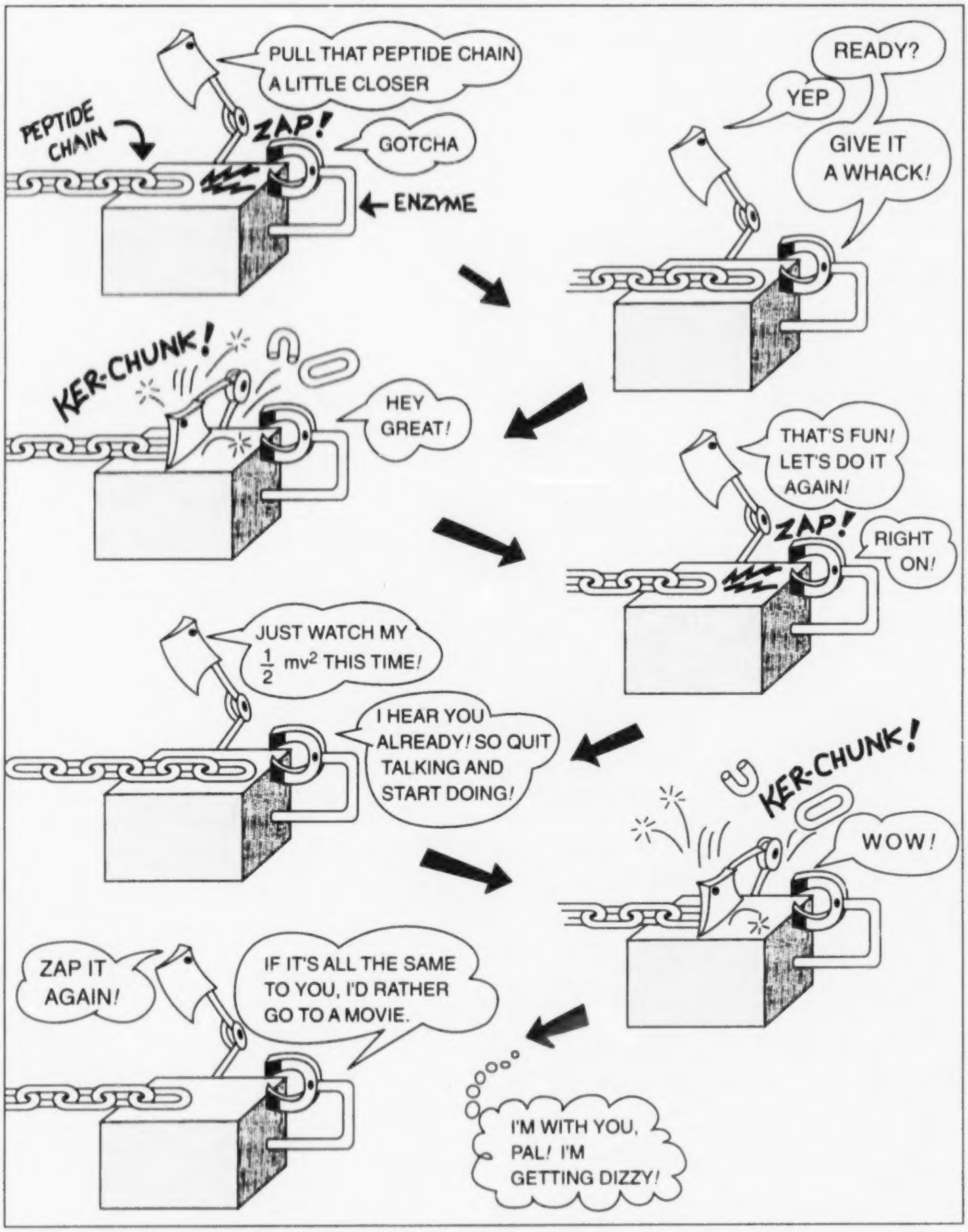
PULL THAT PEPTIDE CHAIN A LITTLE CLOSER
PEPTIDE CHAIN
ZAP!
GOTCHA
ENZYME
READY?
YEP
GIVE IT A WHACK!
KER-CHUNK!
HEY GREAT!
THAT'S FUN! LET'S DO IT AGAIN!
ZAP!
RIGHT ON!
JUST WATCH MY $\frac{1}{2}$ mv^2 THIS TIME!
I HEAR YOU ALREADY! SO QUIT TALKING AND START DOING!
KER-CHUNK!
WOW!
ZAP IT AGAIN!
IF IT'S ALL THE SAME TO YOU, I'D RATHER GO TO A MOVIE.
I'M WITH YOU, PAL! I'M GETTING DIZZY!

2. There's an interesting biochem caper—
A future world food supply shaper.
Seem's bacteria'll toil
To make protein from oil
Or just think—juicy steak from newspaper!

CHEM 13 NEWS, April 1975, p. 894

3. Word Search (see *Appendix* for master copy)

J E ~~Z Y M O G E N~~ O F P K A I U C
L D J D W H M B R H C P L I S H K
Y E I H ~~E D I T P E P~~ P A O O T C
X N K P W G G J T Q H O A P E R E
S A O R M P T V X A W L F D L N L
F T ~~C I R E T O H P M A~~ I S E Z Q
V U W X U K M E K M G F Z N C A N
E R A O A J L U C Y L X Z Y T M W
O A Q J N I N J B U X Y V N R R P
K T J Y X E J R S T M X I Y I N Z
K I E K J S Y I F E H O S P C Y R
O O Z J O A ~~D I C A O N I M A~~ U K
C N ~~S I S E R O H P O R T C E L E~~

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

1. Model of protein structure in which intrachain hydrogen bonding holds polypeptide chains together in a coil. (2 words)
2. Building block of proteins. (2 words)
3. Having both acidic and basic character.
4. The loss of form and function of a protein.
5. Type of bond formed between sulfur atoms of polypeptide chains.
6. Technique used for separating proteins based on their molar mass and charge.
7. pH point at which an amino acid has no net charge.
8. Another name for the amide bond joining two amino acids.
9. Biocatalyst protein.
10. Another name for a proenzyme or inactive form of an enzyme.

Answers: 1. ALPHA HELIX 2. AMINO ACID 3. AMPHOTERIC
4. DENATURATION 5. DISULFIDE 6. ELECTROPHORESIS
7. ISOELECTRIC 8. PEPTIDE 9. ENZYME 10. ZYMOGEN

4. See cartoons at end of module.

MEDIA

1. *The World of Chemistry* videotape "Number 23: Proteins: Structure and Function," is a good review of proteins with emphasis on primary and secondary structure. World of Chemistry Videocassettes. Annenberg/CPB Project, P.O. Box 1922, Santa Barbara, CA 93116-1922; (800) 532-7637; World of Chemistry Series, Atlantic Video, 150 South

Gordon Street, Alexandria, VA 22304; (703) 823-2800 or QUEUE Educational Video, 338 Commerce Drive, Fairfield, CT 06430; (800) 232-2224.

2. The following programs are available from Queue Educational Video, 338 Commerce Drive, Fairfield, CT 06430, (800) 232-2224:
    a. *Digestive Systems.* JNS4048A Apple II's, 64K Disk (Emphasizes digestion. Allows student interaction.)
    b. *The Basis of Biology: Biochemistry.* COM4232A, Apple II's, 64K, Disk and COM4233B, IBM PC, 128K, Disk. (Animated demonstrations of metabolism, including proteins, ATP, coenzymes, RNA, and DNA.)
    c. *Chemistry of Life*, Apple II, IBM-PC, Macintosh. (Tutorial for extensive review of cell chemistry.)
    d. *Enzyme Investigations* (HRM Software), HRM543A Apple II's. (Interactive introduction to enzymes and the digestive system.)
    e. *Biochemistry Unit.* BI04012A, Apple II's. Five disks (Basic biochemistry.)
3. The following programs are available from Cambridge Development Laboratories, Inc., 214 Third Avenue, Waltham, MA 02154.
    a. *Functional Chemistry in Living Cells* (HRM Film/Video) 7-HRC-054F. (A series of videos/films that explore the chemistry of life from the perspectives of cells and cell functions.)
    b. *Proteins* Apple II's. 3-BLB-244A (Tutorial presentation of properties of proteins.)
4. Software published by Project SERAPHIM, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 263-2837 (voice) or (608) 262-0381 (FAX).
    a. For the Apple II computer: AP 726 (*Protein Structure Prediction* and *Restriction Map Simulator*)
    b. For IBM PCs and PC-compatibles: PC 4901 (same as AP 726)
5. Videodisc published by *JCE: Software*, a publication of the *Journal of Chemical Education*, Department of Chemistry, University of Wisconsin-Madison, 1101 University Avenue. Madison, WI 53706-1396: (608) 262-5153 (voice) or (608) 262-0381 (FAX).

    "The Boat and Dock Model," a chapter on *The World of Chemistry: Selected Demonstrations and Animations:* Disc II (double sided, 60 min.), Special Issue 4.

INSTRUMENTATION

1. Vernier's Voltage Plotter with pH electrode. Price: approx $100.
2. Nester model and Corning model pH meters (Flinn Scientific, Inc. Chemical Catalog/Reference Manual, 131 Flinn Street, P. O. Box 219, Batavia, IL 60510-0219). Price range $25-700.
3. The following enzyme kits are available from Flinn Scientific Co., P.O. Box 219, 131 Flinn Street, Batavia, IL 60510.
    a. *Enzymes: A Qualitative Approach* AP 1771 (Classroom quantities of materials needed to show how invertase and protease activity may be detected.)
    b. *Biotechnology Enzymes in the Home* AP1770 (Kit of materials to show industrial applications of certain enzymes.)
4. The following instructional kits are available from Fisher Scientific Co., Educational Materials, Division, 4901 West LeMoyne Street, Chicago, IL 60651.
    a. Chromatography/Electrophoresis Kit S40085ND (All apparatus and materials for 41 activities for separating materials by paper chromatography and electrophoresis.)
    b. Electrophoresis Kit A14077. (Apparatus for separation of proteins and amino acids. Requires power supply unit sold separately [S40087ND].)

Links/Connections

WITHIN CHEMISTRY

1. Chemical Kinetics
2. Acid/Base
3. Biochemistry (proteins)
4. Stereochemistry (isomers)

BETWEEN CHEMISTRY AND OTHER DISCIPLINES

1. **Nutrition**. As enzyme activators, vitamins and minerals are dietary essentials.
2. **Medicine**. Abnormalities often result if specific enzymes are missing. Over one hundred disorders due to deficient activity of a specific enzyme have been demonstrated in humans (see text by Holum in *References*).
3. **Pest Control**. Studies are under way to identify the enzyme that allows the cockroach to synthesize its exoskeleton. Since the roach can not live without this crusty shell, inactivating this enzyme will help eradicate this pest.

TO THE CONTEMPORARY WORLD

1. Fermentation results because of enzymatic action.
2. The enzyme *streptokinase* is often injected into the blood stream of patients who have suffered a heart attack from clots in the coronary arteries. The enzyme often dissolves the clots and saves the patient.
3. *Proteinase* is a general term for enzymes that helps dissolve spots and stains on clothing. They are used extensively in laundry detergents and are produced in larger quantities than any other commercial enzyme.
4. Many enzymes are used in food processing, including:
    a. **Amylase**: Breaks down starch to produce more sugar for fermentation. Used extensively in bakery products, brewing, and production of syrups, sugars, and cereals.
    b. **Catalase**: Used in the "cold pasteurization" of milk.
    c. **Cellulase**: Helps break down cellulose in coffee beans.
    d. **Invertase**: Converts sucrose to glucose and fructose for a sweeter taste. Used in honey and candies.
    e. **Pectinase**: Used to break down cell structures in fruits, increasing the amount of juice (also applesauce from apples).
    f. **Rennet**: Coagulates casein in milk, producing curd.
5. Sewage Treatment. Both amylase and proteases are used to break down insoluble carbohydrates and proteins in sewage.
6. Amylases are used in the paper industry as binding agents for starch and other fillers used to make paper. They are also used to desize fabrics in the textile industry.
7. Proteases are used to prepare leather by removing hair and other material from hides. They are also used in the photography industry to reclaim silver from used film and developer.

Extensions

1. Have students do library research on enzymes involved in digestion. For example, *amylase* catalyzes the breakdown of starch; *lipases* catalyze the breakdown of fats, *etc.*
2. Have students talk with physicians about *Inborn Errors of Metabolism*. Two common examples are phenylketonuria and galactosemia.
3. Many states have laws that require the screening for certain enzyme defects in the newborn. Have students visit state or local health departments to learn about their policies.
4. Many animals (horses, cows, goats, termites, *etc.*) have bacteria in their intestinal system that produce the enzyme *cellulase*. This enzyme breaks down cellulose and permits these animals to digest grass, wood, paper, *etc.* Unfortunately, humans do not have the bacterium that produces this enzyme. Have students discuss the theoretical ramification if this enzyme could somehow be made available to humans.
5. All vitamins (except A, D, E, K) serve as enzyme *activators* in biochemical reactions. Have students do library research to learn more about the specific role of each vitamin.
6. Try gelatin (Jello™) prepared with fresh pineapple and also with canned pineapple. Why does the preparation made with fresh pineapple not gel?
7. Have students prepare reports on the importance of enzymes in cooking or home products.

References

Module developed by Robert Sol Davis, Patricia Owens, and Lee Summerlin, the Alabama team.

Borgford, C. L., and Summerlin, L. R. (1988). *Chemical activities.* Washington, DC: American Chemical Society.

Holum, J. R. (1990). *Fundamentals of general, organic, and biological chemistry* (4th Ed.). New York, NY: John Wiley & Sons.

Chapter 24 contains a good description of enzymes, how they work, and their usefulness in medicine.

Lehninger, A. L. (1988). *Biochemistry.* New York, NY: Worth.

Novo Laboratories, Inc. (1975). *Enzymes, nature's catalysts.* Burlington, NC: Carolina Biological Supply.

Ornstein, A. (September, 1993). Alcohol denatures proteins, *Chem 13 News*, p. 31.

Summerlin, L. R. (1981). *Chemistry for the life sciences.* New York, NY: Random House.

Summerlin, L. R. (1979). *Chemistry of common substances.* Morristown, NJ: Silver-Burdett.

Summerlin, L. R., Borgford, C. L., and Ealy, J. B. (1988). *Chemical demonstrations: A sourcebook for teachers* (Vol. 2, 2nd Ed.). Washington, DC: American Chemical Society.

Appendix

- **Transparency Master**
  Word Search
- **Humor**

Word Search

J E Z Y M O G E N O F P K A I U C
L D J D W H M B R H C P L I S H K
Y E I H E D I T P E P P A O O T C
X N K P W G G J T Q H O A P E R E
S A O R M P T V X A W L F D L N L
F T C I R E T O H P M A I S E Z Q
V U W X U K M E K M G F Z N C A N
E R A O A J L U C Y L X Z Y T M W
O A Q J N I N J B U X Y V N R R P
K T J Y X E J R S T M X I Y I N Z
K I E K J S Y I F E H O S P C Y R
O O Z J O A D I C A O N I M A U K
C N S I S E R O H P O R T C E L E

Words about the concepts in this module can be obtained from the clues given. Find these words in the block of letters:

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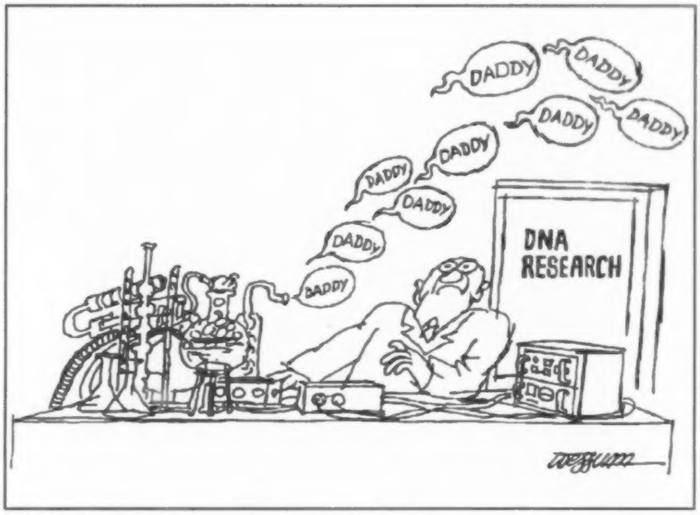
7. pH point at which an amino acid has no net charge.
8. Another name for the amide bond joining two amino acids.
9. Biocatalyst protein.
10. Another name for a proenzyme or inactive form of an enzyme.

SANDY
"But is it right to bring a recombinant being into a world like this?"

CHEMTECH, August 1984, p. 485. ©American Chemical Society. Reprinted with permission.


DADDY
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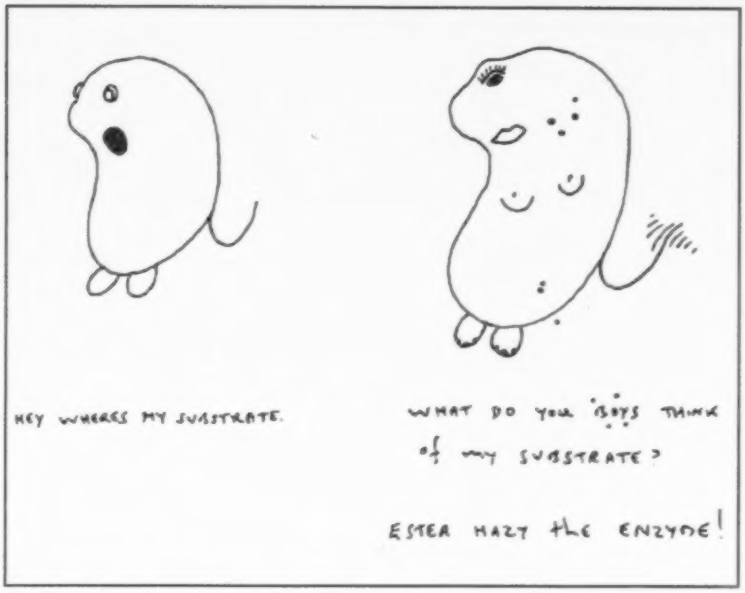
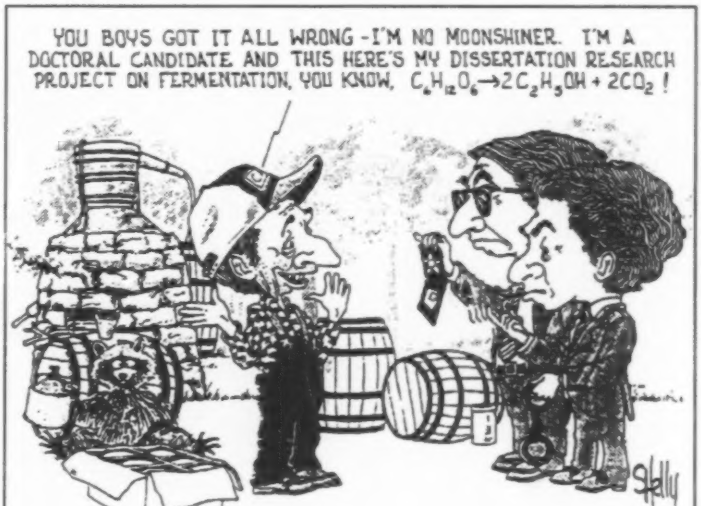
DNA RESEARCH

CHEMTECH, October 1984, p. 578. ©American Chemical Society. Reprinted with permission.


HEY WHERES MY SUBSTRATE.
WHAT DO YOU BOYS THINK of MY SUBSTRATE?
ESTER HAZY the ENZYME!

CHEM 13 NEWS, October 1971, p. 341. Reprinted with permission.


YOU BOYS GOT IT ALL WRONG - I'M NO MOONSHINER. I'M A DOCTORAL CANDIDATE AND THIS HERE'S MY DISSERTATION RESEARCH PROJECT ON FERMENTATION, YOU KNOW, $C_6H_{12}O_6 \rightarrow 2C_2H_5OH + 2CO_2$!
Shelly

Cartoon by Shelly Fischman. Reprinted with permission.

JAN-21-97 TUE 13:54

P.04

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College of New Rochelle

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